Name: ______________________________

Garden Grove High School

Chemistry

Notes and
Worksheets for Units
1-5

2024-25

Units of Chemistry

Semester 1:

- Unit 1: Class Intro and Laboratory Safety
- Unit 2: Properties of Matter
- Unit 3: Atomic Structure and the Periodic Table
- Unit 4: Modern Atomic Theory
- Unit 5: Chemical Bonding and Intermolecular Forces
- Unit 6: Chemical Naming
- Unit 7: Understanding and Predicting Chemical Reactions
- Unit 8: Scientific Measurement and Calculation (begin)

Semester 2:

- Unit 8: Scientific Measurement and Calculation (finish)
- Unit 9: Energy in Chemical Reactions
- Unit 10: The Mole and Chemical Composition
- Unit 11: Stoichiometry
- Unit 12: Properties of Aqueous Solutions
- Unit 13: Acids and Bases
- Unit 14: Chemical Equilibrium

Periodic Table of the Elements

1 **H** Hydrogen 1.008																	2 **He** Helium 4.003
3 **Li** Lithium 6.941	4 **Be** Beryllium 9.012											5 **B** Boron 10.811	6 **C** Carbon 12.011	7 **N** Nitrogen 14.007	8 **O** Oxygen 15.999	9 **F** Fluorine 18.998	10 **Ne** Neon 20.180
11 **Na** Sodium 22.990	12 **Mg** Magnesium 24.305											13 **Al** Aluminum 26.982	14 **Si** Silicon 28.086	15 **P** Phosphorus 30.974	16 **S** Sulfur 32.066	17 **Cl** Chlorine 35.453	18 **Ar** Argon 39.948
19 **K** Potassium 39.098	20 **Ca** Calcium 40.078	21 **Sc** Scandium 44.956	22 **Ti** Titanium 47.867	23 **V** Vanadium 50.942	24 **Cr** Chromium 51.996	25 **Mn** Manganese 54.938	26 **Fe** Iron 55.845	27 **Co** Cobalt 58.933	28 **Ni** Nickel 58.693	29 **Cu** Copper 63.546	30 **Zn** Zinc 65.38	31 **Ga** Gallium 69.723	32 **Ge** Germanium 72.631	33 **As** Arsenic 74.922	34 **Se** Selenium 78.971	35 **Br** Bromine 79.904	36 **Kr** Krypton 83.798
37 **Rb** Rubidium 85.468	38 **Sr** Strontium 87.62	39 **Y** Yttrium 88.906	40 **Zr** Zirconium 91.224	41 **Nb** Niobium 92.906	42 **Mo** Molybdenum 95.95	43 **Tc** Technetium 98.907	44 **Ru** Ruthenium 101.07	45 **Rh** Rhodium 102.906	46 **Pd** Palladium 106.42	47 **Ag** Silver 107.868	48 **Cd** Cadmium 112.414	49 **In** Indium 114.818	50 **Sn** Tin 118.711	51 **Sb** Antimony 121.760	52 **Te** Tellurium 127.6	53 **I** Iodine 126.904	54 **Xe** Xenon 131.294
55 **Cs** Cesium 132.905	56 **Ba** Barium 137.328	57-71	72 **Hf** Hafnium 178.49	73 **Ta** Tantalum 180.948	74 **W** Tungsten 183.84	75 **Re** Rhenium 186.207	76 **Os** Osmium 190.23	77 **Ir** Iridium 192.217	78 **Pt** Platinum 195.085	79 **Au** Gold 196.967	80 **Hg** Mercury 200.592	81 **Tl** Thallium 204.383	82 **Pb** Lead 207.2	83 **Bi** Bismuth 208.980	84 **Po** Polonium [208.982]	85 **At** Astatine 209.987	86 **Rn** Radon 222.018
87 **Fr** Francium 223.020	88 **Ra** Radium 226.025	89-103	104 **Rf** Rutherfordium [261]	105 **Db** Dubnium [262]	106 **Sg** Seaborgium [266]	107 **Bh** Bohrium [264]	108 **Hs** Hassium [269]	109 **Mt** Meitnerium [278]	110 **Ds** Darmstadtium [281]	111 **Rg** Roentgenium [280]	112 **Cn** Copernicium [285]	113 **Nh** Nihonium [286]	114 **Fl** Flerovium [289]	115 **Mc** Moscovium [289]	116 **Lv** Livermorium [293]	117 **Ts** Tennessine [294]	118 **Og** Oganesson [294]
	Lanthanide Series	57 **La** Lanthanum 138.905	58 **Ce** Cerium 140.116	59 **Pr** Praseodymium 140.908	60 **Nd** Neodymium 144.243	61 **Pm** Promethium 144.913	62 **Sm** Samarium 150.36	63 **Eu** Europium 151.964	64 **Gd** Gadolinium 157.25	65 **Tb** Terbium 158.925	66 **Dy** Dysprosium 162.500	67 **Ho** Holmium 164.930	68 **Er** Erbium 167.259	69 **Tm** Thulium 168.934	70 **Yb** Ytterbium 173.055	71 **Lu** Lutetium 174.967	
	Actinide Series	89 **Ac** Actinium 227.028	90 **Th** Thorium 232.038	91 **Pa** Protactinium 231.036	92 **U** Uranium 238.029	93 **Np** Neptunium 237.048	94 **Pu** Plutonium 244.064	95 **Am** Americium 243.061	96 **Cm** Curium 247.070	97 **Bk** Berkelium 247.070	98 **Cf** Californium 251.080	99 **Es** Einsteinium [254]	100 **Fm** Fermium 257.095	101 **Md** Mendelevium 258.1	102 **No** Nobelium 259.101	103 **Lr** Lawrencium [262]	

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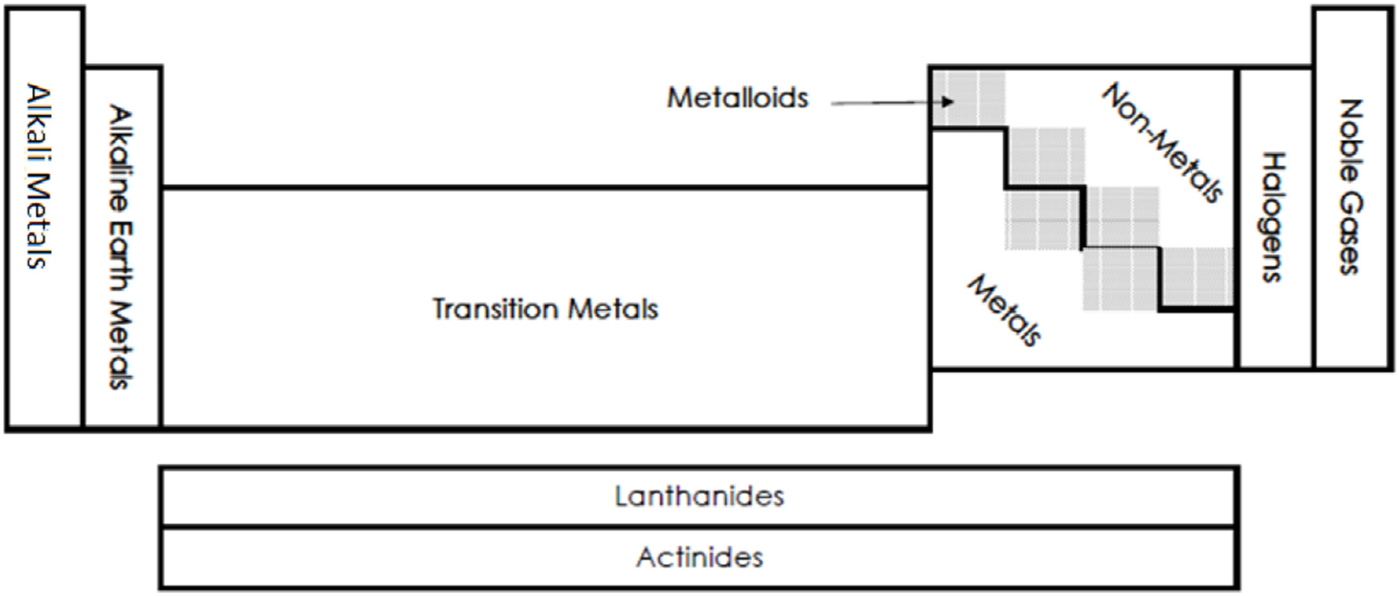
Regions of the Periodic Table



KNOW THE FOLLOWING ELEMENT NAMES AND SYMBOLS:

NAME	SYMBOL		NAME	SYMBOL		NAME	SYMBOL
aluminum	Al		chromium	Cr		manganese	Mn
argon	Ar		sulfur	S		mercury	Hg
arsenic	As		tin	Sn		nickel	Ni
barium	Ba		fluorine	F		nitrogen	N
beryllium	Be		gold	Au		oxygen	O
boron	B		helium	He		phosphorus	P
bromine	Br		hydrogen	H		potassium	K
calcium	Ca		iodine	I		silicon	Si
carbon	C		iron	Fe		silver	Ag
chlorine	Cl		lead	Pb		sodium	Na
cobalt	Co		lithium	Li		strontium	Sr
copper	Cu		magnesium	Mg		zinc	Zn

COMMON POLYATOMIC IONS			
Ion	Name	Ion	Name
NH_4^+	Ammonium	HCO_3^-	Hydrogen carbonate
NO_2^-	Nitrite		(bicarbonate)
NO_3^-	Nitrate	BrO_3^-	Bromate
SO_3^{2-}	Sulfite	CO_3^{2-}	Carbonate
SO_4^{2-}	Sulfate	ClO^-	Hypochlorite
HSO_4^-	Hydrogen sulfate	ClO_2^-	Chlorite
	(bisulfate)	ClO_3^-	Chlorate
OH^-	Hydroxide	ClO_4^-	Perchlorate
CN^-	Cyanide	$C_2H_3O_2^-$	Acetate
PO_4^{3-}	Phosphate	MnO_4^-	Permanganate
HPO_4^{2-}	Hydrogen phosphate	$Cr_2O_7^{2-}$	Dichromate
HSO_3^-	Hydrogen sulfite	CrO_4^{2-}	Chromate
H_3O^+	Hydronium	$C_2O_4^{2-}$	Oxalate

International System of Units (SI)

SI Base Units

Base Quantity	Name	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electric current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

SI Derived Units

Derived Quantity	Name	Symbol	Equivalent SI units
Frequency	hertz	Hz	s^{-1}
Force	newton	N	$m{\cdot}kg{\cdot}s^{-2}$
Pressure	pascal	Pa	N/m^2
Energy	joule	J	N·m
Power	watt	W	J/s
Electric charge	coulomb	C	s·A
Electric potential	volt	V	W/A
Electric resistance	ohm	Ω	V/A
Celsius temperature	degree Celsius	°C	K*
Volume	liter	L	dm^3

SI Prefixes

Factor	Name	Symbol	Numerical Value
10^{12}	tera	T	1 000 000 000 000
10^{9}	giga	G	1 000 000 000
10^{6}	mega	M	1 000 000
10^{3}	kilo	k	1 000
10^{2}	hecto	h	100
10^{1}	deka	da	10
10^{-1}	deci	d	0.1
10^{-2}	centi	c	0.01
10^{-3}	milli	m	0.001
10^{-6}	micro	µ	0.000 001
10^{-9}	nano	n	0.000 000 001
10^{-12}	pico	p	0.000 000 000 001

- Adapted from NIST Special Publication 811
- SI rules and style conventions recommend using spaces rather than commas to separate groups of three digits.

ACADEMY FOR MATH, ENGINEERING & SCIENCE

METRIC PREFIX	SYMBOL	NUMERICAL EQUIVALENT		
yotta	(Y)	10^{24}	1 septillion	1,000,000,000,000,000,000,000,000
zetta	(Z)	10^{21}	1 sextillion	1,000,000,000,000,000,000,000
exa	(E)	10^{18}	1 quintillion	1,000,000,000,000,000,000
peta	(P)	10^{15}	1 quadrillion	1,000,000,000,000,000
tera	(T)	10^{12}	1 trillion	1,000,000,000,000
giga	(G)	10^{9}	1 billion	1,000,000,000
mega	(M)	10^{6}	1 million	1,000,000
kilo	(k)	10^{3}	1 thousand	1000
hecto	(h)	10^{2}	1 hundred	100
deka	(da)	10	1 ten	10
deci	(d)	10^{-1}	1 tenth	0.1
centi	(c)	10^{-2}	1 hundredth	0.01
milli	(m)	10^{-3}	1 thousandth	0.001
micro	(µ)	10^{-6}	1 millionth	0.000001
nano	(n)	10^{-9}	1 billionth	0.000000001
pico	(p)	10^{-12}	1 trillionth	0.000000000001
femto	(f)	10^{-15}	1 quadrillionth	0.000000000000001
atto	(a)	10^{-18}	1 quintillionth	0.000000000000000001
zepto	(z)	10^{-21}	1 sextillionth	0.000000000000000000001
yocto	(y)	10^{-24}	1 septillionth	0.000000000000000000000001

Other Useful Equivalents

Length

1Å (Ångstrom) = 1.0×10^{-10} m

12 in = 1 ft

1 mile = 5,280 ft

1 mile = 1,760 yards

Volume

1 L = 10^{-3} m^3

1 L = 1 dm^3

1 L = 10^3 cm^3

1 mL = 1 cm^3

1 gal = 4 quarts

1 gal = 8 pints

1 gal = 128 fl oz

1 quart = 2 pints

Mass

1 ton (USA only)

= 2,000 lbs

1 metric ton

= 1,000 kg

Temperature

K = °C + 273

CONVERSION FACTORS

LENGTH

1 m = 39.37 in = 3.28 ft

1 inch = 2.54 cm

1 cm = 0.394 in

1 ft = 0.305 m

1 km = 0.621 mi

1 mi = 1.61 km

MASS/WEIGHT

1 kg = 2.2046 lb*

1 lb = 0.454 kg

1 g = 0.0353 oz

1 oz = 28.35 g

** A kilogram is a unit of mass while a pound is a unit of weight which is mass under the influence of gravity. This makes it a force. This conversion might work on the surface of the Earth but not anywhere else in the Universe.*

TEMPERATURE

Degrees F = (9/5)x+32 *(where x=degrees C)*

Degrees F = (9/5)x-459.67 *(where x=degrees K)*

Degrees C = (5/9)(x-32) *(where x=degrees F)*

Degrees C = x+273.15 *(where x=degrees K)*

Degrees K = x-273.15 *(where x=degrees C)*

Degrees K = (9/5)x-459.67 *(where x=degrees F)*

VOLUME

1 L = 0.2642 gal

1 L = 33.81 fl oz

1 gal = 3.7854 L

1 gal = 128 fl oz

1 mL = 1 cc (cm^3)

1 fl oz = 29.574 mL

1 mL = 0.034 fl oz

Non-SI Units often used in Chemistry

Quantity	Non-SI Unit Name	Non-SI Unit Symbol	SI Unit Name	SI Unit Symbol
Volume	Liter Milliliter	L mL	Cubic meter	m^3
Density	Gram/milliliter	g/mL	Kilogram per cubic meter	kg/m^3
Temperature	Degree Celsius	°C	Kelvin	K
Energy	calorie	cal	Joule	J
Pressure	Atmosphere of pressure, Milliliter of mercury (also a measure of pressure)	atm mmHg	Pascal	Pa

Quantity	Name	Symbol	SI base unit	Derived unit
Area	square metre	A	m^2	-
Volume	cubic metre	V	m^3	-
Velocity	metre per second	v	m/s	-
Acceleration	metre per second squared	a	m/s^2	-
Force	newton	N	$kg\ m/s^2$	J/m
Energy	joule	J	$kg\ m^2/s^2$	N m
Pressure or stress	pascal	Pa	$kg\ m/s^2$	N/m^2
Power	watt	W	$kg\ m^2/s^3$	J/s

Unit 1: Class Introduction and Safety in the Chemistry Laboratory

Name______________________________ Per: _____

Classroom Policies for Chemistry

GRADES

Student work is counted in weighted categories as follows:

- 15% Homework/classwork – labs and assignments or worksheets completed in class and/or at home
- 20% Formative assessments – quizzes
- 50% Summative – unit exams
- 15% Final exam – covers the entire semester

GRADING SCALE (Grades will NOT be rounded up, however + and - grades are given within each letter range)

A: 90 – 100.00% (above 100% is an A+)
B: 80.00 – 89.99%
C: 70.00 – 79.99%
D: 60.00 – 69.99%
F: 0 – 59.99%

COLLEGE ELIGIBILITY INFORMATION

Completion of this course with a grade of C- or better satisfies the second year of the two-year D (Science) requirement of the A-G requirements for UC and CSU eligibility. The minimum is 2 years in D (Science), but due to the competitive nature of college admissions, it is strongly recommended to complete four, or at the least, three years in science.

WEBSITE – Many assignments and class materials will be posted on the class website. Typing in either of these two URLs will get you to the website: **holguinscience.com** or **holguinscience.weebly.com**

GENERAL RULES

- **EXTRA CREDIT** – Although the instructor might, on rare occasion, choose to give special extra credit assignments, but do not expect them regularly. Extra credit may, however (at teacher discretion) be given for things like early completion of assignments.

- **COMPLETENESS/EFFORT/NEATNESS** – All assignments must be complete in order to receive a grade. Incomplete work gets a score of zero. If directions are not followed and/or work is done incorrectly, then the work is considered incomplete. Any assignment that is not easily readable or lacks evidence of effort will not receive credit.

- **CHROMEBOOKS** – Students should come to school with their chromebook (and keep it fully charged), and should have it with them every day. Students should also bring a chromebook charger with them in case the battery runs low. It is the responsibility of the student to contact the school and get the chromebook repaired or replaced if needed.

- **CLASSROOM MATERIALS** – At a minimum, students should have the following materials when in class: notebook (5-subject, college-ruled, size 8.5 × 11 inches or 9 × 11 inches), textbook, scientific calculator (programmable calculators are not allowed on quizzes or exams) paper, and pencil or pen. Notebooks will be turned in (for the teacher to keep) at the end of the school year. Recommended calculator: TI-30S IIX

- **ACADEMIC DISHONESTY (CHEATING/COPYING) – DON'T DO IT!!!**
    - Students caught cheating will receive an automatic "no credit" for the assignment, quiz or test and may be referred to school administration, where it will be noted on their permanent school record.
    - Repeat offenders can be given a grade of "F" for the course (meaning that they have to redo the entire class)
    - Examples of cheating include but are not limited to:
        - Turning in any work that is not your own while claiming or implying that it is yours
        - Getting unauthorized help on a test/quiz
        - Telling people about what is on a quiz or test that they haven't taken yet, or receiving information about one that you have not yet taken

- **EXAMS** – Unless otherwise noted, tests and quizzes are taken without books or notes. A cumulative final exam will be given at the end of each semester. Study guides are not given, so plan on making your own!

- **LAB** – In order to participate in lab, students MUST dress appropriately, and MUST have completed any and all pre-lab assignment(s). Use of safety equipment and observance of safe behavior and procedures is also required, and students who violate these expectations will be excluded from lab. Students arriving late on lab day may also be excluded from lab. Students must bring lab assignments with them to lab in order to participate. Student excluded from lab for any reason receive no credit for the lab.

- **NOTES** – Students are required to take notes and should have an appropriate notebook in order to do so. Students should also maintain an accurate table of contents. Some notes will need to be completed (and studied) outside class hours.

- **HOMEWORK** – In order to get credit, all homework must be complete, easily readable, show all your work, and have a boxed or highlighted answer. Chemistry is like math class in many ways, and if work is not shown, then no credit is given, even if the answer is correct. It is important to complete the homework, as it is the practice that allows you to learn the material well and score better on the exams.

- **LATE WORK** – In general, homework should be done outside of class. Homework not completed by the start of class or not turned in when called for will be considered late. All homework handed in by students who arrive tardy is also late. If the teacher decides to accept late work (this is not a guarantee), then it might get up to 50% credit up to one week after the due date or up until the day before the unit exam, whichever is first. Students who miss an assignment, quiz or exam due to arriving late to class receive a score of zero for that assignment and will not generally be allowed extra time or the opportunity to make up the lost points.

- **ABSENT WORK POLICY** – Students with excused absences will receive a reasonable amount of extra time to complete work in keeping with the time actually needed to reasonably do so. Digital work, such as edpuzzles, google classroom assignments, etc can easily be done when absent and are not eligible for extra-time accommodations except by prior request and teacher approval. Paper-based in-class work assigned when a student is gone generally comes with reasonable extra time to complete the work, e.g. equal to the number of days they were absent if the absence occurred on the day it was assigned. If a student is absent on a due date for paper-based work, then all work due that day must be turned in digitally at the soonest possibility, or in person immediately upon return to school. Students with truancies or unexcused absences have no right to any kind credit for missed assignments, labs, quizzes, exams, etc.

- **BATHROOM/HALL PASSES** – Students may have 1 bathroom/hall pass per quarter. Those who do not use their hall pass will receive extra credit. Students with no passes who need to leave the classroom will have points deducted from their class grade each time. Students absent when hall passes are given out (typically first day of the quarter) do not get any if they don't ask for one upon returning.

- **EXTRA HELP OUTSIDE OF CLASS** – Your teacher is generally available for help outside class hours, so ask them which hours are best. If your teacher is not available, then any of the other chemistry teachers (see list below) can also help. Aside from your instructor, homework help can also be had through AVID tutoring and various other before- and after-school tutoring programs.

CLASS RULES

Respect Class Time

- Be on time
- Start work promptly.
- Bring all materials to class.
- Ask for help

Respect Others

- Cooperate.
- Listen.

Respect the Room

- Use materials and supplies correctly.
- Clean up after yourself.
- Keep food and drink outside.

In addition to the rules outlined here, all school-wide rules will be STRICTLY enforced.

CONSEQUENCES

Students who choose not to follow these expectations may earn one or more of the following consequences:

- Warning
- Movement to another seat
- Time out outside of the room
- After class counseling (detention)
- Referral for administrative discipline
- Parent contact

Any student engaging in chronic or severe misbehavior will be referred to the appropriate administrator for discipline.

If you have any questions, contact your teacher at the appropriate email address. If you have questions or need help with chemistry-related issues, then any of these teachers can be of assistance:

- Mr Holguin – cholguin@ggusd.net
- Mr Luong – vluong@ggusd.net
- Ms Lynch – blynch@ggusd.net
- Mr Oda – joda@ggusd.net

Review this syllabus with your parents and return to class as soon as possible.

By signing, the student agrees to the following statement: "I have read this syllabus and understand that I am expected to contribute to and participate in this class in a manner that is responsible, respectful and productive."

Student signature: ______________________ Student print name: ______________________ Date: ________

Parent signature: ______________________ Parent print name: ______________________ Date: ________

Science Lab Safety Rules

Remember that violation of any of the rules on this sheet is grounds for immediate expulsion from lab with a grade of zero.

1. Always wear goggles, closed-toe shoes, and appropriate clothing. No exceptions.

2. Conduct yourself in a responsible manner at all times in the laboratory.

3. Study your lab assignment before you come to class and have your prelab completed. Pay particular attention to dangerous reagents or special safety precautions.

4. Never work in the lab without proper supervision.

5. Never perform an unauthorized experiment.

6. Never inhale gases or other vapors.

7. Be careful with glass equipment: hot glass looks just like cool glass!

8. Never chew gum, eat or drink or in any part of the lab classroom

9. Report any accident, injury, chemical spill or inappropriate procedure at once.

10. Tie back long hair and secure loose clothing.

11. Do not leave a lighted burner unattended.

12. Keep your lab area as clean as possible at all times.

13. Never return unused chemicals to a reagent bottle

14. Do not allowing pieces of equipment to come in contact with each other, this causes cross-contamination

15. Do not handle dangerous substances or concentrated acids unless instructed to do so by your teacher. In general, it is best to leave this sort of task to those who have experience in handling these sorts of materials, since they are likely to be aware of hazards that you don't know about.

16. Place broken glass, solid substances and used chemicals in designated container. Never pour or wash a chemical down the drain unless told to do so by your instructor

17. Wash your hands with soap and water after each experiment.

18. To test the odor of a solution, do not hold it directly under nose. Hold it away and wave your hand over the solution towards you (wafting)

Science Laboratory Safety Contract - GGHS Chemistry

1. I will act responsibly at all times in the laboratory.
2. I will wear my safety goggles, closed-toe shoes, and appropriate clothing at all times, and understand that failure to do this can result is expulsion from lab with a grade of zero with NO WARNING.
3. I will complete all pre-lab assignments and bring my lab papers with me on lab day.
4. I will follow all lab safety rules when in the classroom.
5. I will follow all oral and written instructions provided by my teacher.
6. I will immediately notify the teacher of an emergency.
7. I will tie back long hair and remove jewelry while in the laboratory and wear close toed shoes when necessary.
8. I will not take chemicals or equipment out of the laboratory unless instructed to do so by the teacher.
9. I will never eat or drink in the laboratory.
10. I will not enter the chemical storage room at any time.
11. I know the locations of the fire extinguisher, fire blanket, shower and eyewash station in the classroom.
12. I will keep my area in the laboratory clean.
13. This contract will stay in my science notebook at all times.

I, ____________________________ (student's printed name) have been instructed in the lab safety and emergency precautions needed for my science class. I have read and understand the lab safety rules on the other side of this paper and the science laboratory safety contract. I am aware that my safety and the safety of my classmates depend on my behavior in the laboratory. I agree to follow all safety rules and any additional instructions provided by my teacher and/or school administrator.

_________ (date) ___________________________ (student signature)

_________ (date) ___________________________ (parent/guardian signature)

[Insert Advice from Last Year’s Chemistry Students Here]

Homework Checklist

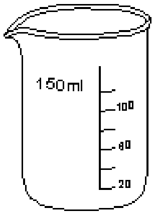
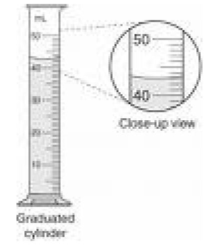
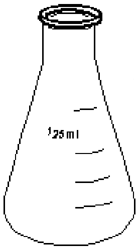
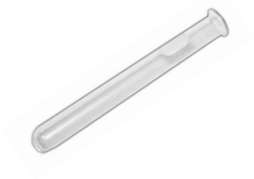
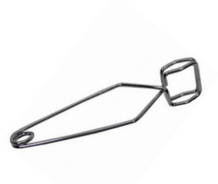
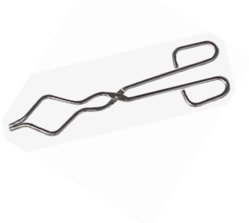
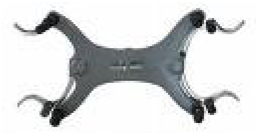
It’s easy to miss a step sometimes, so here’s a little something that might help! Review this list before turning in work, because following these steps is an easy way to get an "A" on all of your assignments!

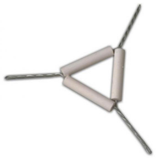
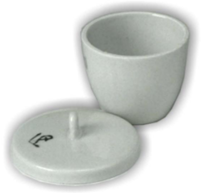
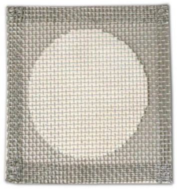
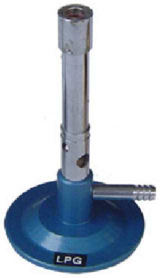
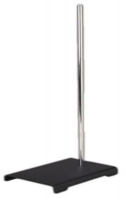
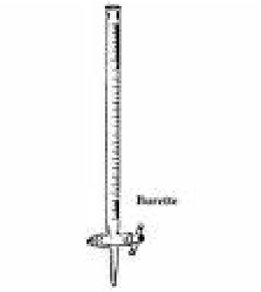
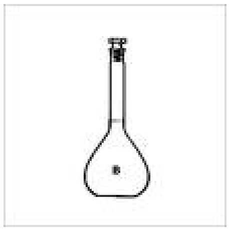
Be sure that you…

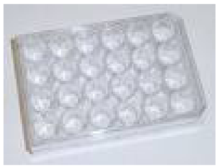
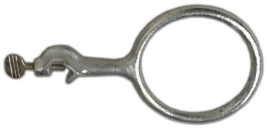
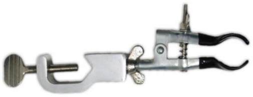
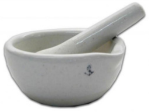
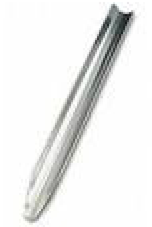
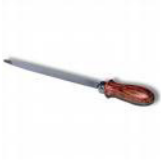
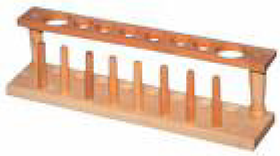
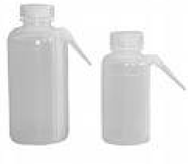
1. Complete the entire assignment (no credit for incomplete work)
2. Obey all written instructions in the assignment
3. Follow any additional oral directions from your instructor
4. Highlight questions you aren’t sure about (minimum 1)
    a. Make sure you highlight entire question rather than just the letter or number
    b. Avoid highlighting entire data/answer tables or large non-specific areas
5. Name (first and last) and period should be written on all work turned in
6. Show work for all calculations
7. Include units for all numbers unless otherwise specified
8. Box the answer for all calculations
9. Obey all formatting instructions
10. Ensure clarity, neatness and easy readability
11. Answers should be factually correct (to the best of your ability) and show effort
12. Spell correctly (to the best of your ability)

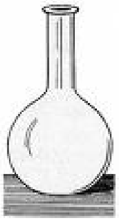
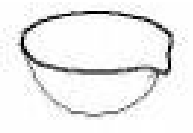
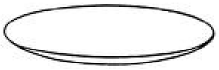
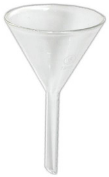
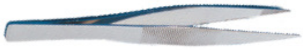
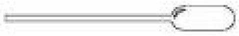
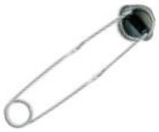
Additional considerations for digital assignments:

13. Super/subscripts are required where needed
14. Check the directions in google classroom
15. Check google classroom announcements to see if there are any updates before completing and submitting
16. Delete any blanks and replace them with answer(s)
17. Respond to questions using a font color different from the original
18. If turning in more than one photo, then it should be
    a. right-side up
    b. inserted into a google doc

Description/Use	Apparatus
Beaker Used for heating liquids, pouring liquids into other containers, mixing (when splashing is not a concern) or measuring volume with low accuracy	
Graduated Cylinder Used to measure volume with Medium to high accuracy and pour liquids into other containers	
Erlenmeyer Flask Glass container used to mix liquids without splashing, heat liquids, or measure volume with low accuracy	
Test Tube Glass container with many uses, can be heated	
Test Tube Holder Metal clamp with a spring handle, Used only for holding test tubes	
Tongs Metal tool used to pick up and hold a variety of high-temperature objects a safe distance away from your hands	
Double Burette Clamp Metal clamp with flexible clips, used to hold burettes	

Description/Use	Apparatus
Pipestem Triangle Triangular wire frame with clay material coverings, used to support the crucible	
Crucible and Cover Small ceramic or metal dish with cover, used to heat small amounts of solid material to very high temperature	
Wire Gauze Wire screen used to spread the heat of a flame from a gas burner	
Gas Burner Heating device connected to a gas outlet by rubber tubing. Used to heat chemicals in ceramic or glass containers	
Ring Stand Upright metal rod with heavy base that is used as a support for lab equipment	
Burette Glass tube used to measure and dispense small volumes of liquids with very high accuracy	
Volumetric Flask Glass container used to measure large volumes of liquids with high accuracy. Has flat bottom.	

Description/Use	Apparatus
Well Plate Plastic with built-in wells, used to hold many small liquid samples	
Support Ring Metal ring with screw fastener, several sizes, is fastened to the ring stand as a support for lab equipment	
Clamp for Test Tubes or Burettes Metal clamp with screw fastener, swivel and lock nut, adjusting screw, and curved clamp, used to hold apparatus, may be fastened to the ring stand	
Mortar and Pestle Heavy porcelain dish with grinder, used to grind chemicals into a powder	
Scoopula Used for transferring solid chemicals to another container for weighing	
Triangular File Metal file with three cutting edges, used to scratch glass to file	
Test Tube Rack Wood, metal, or plastic rack used to hold test tubes in an upright position	
Wash Bottles Various sizes, squeezable plastic bottle with angular tip used to dispense distilled water	

Description/Use	Apparatus
Florence Flask Glass, round-bottomed container used to heat liquids	
Stirring Rod Used to stir chemicals or help remove materials from containers	
Evaporating Dish Porcelain dish used as a container for small amounts of liquid being evaporated	
Watch Glass Curved glass used as a beaker cover or for evaporating small amounts of liquid	
Funnel Glass or plastic used to hold filter paper or used to aid in pouring	
Forceps Tweezers used to pick up or hold small objects	
Thin Stem Disposable Pipette Plastic dropper used to transfer small amounts of liquid	
Flint Striker Metal with flint used to make a spark for lighting flames	

Name: ______________________________ Per: ____

Activity: Safety in the Chemistry Laboratory

Part 1: Understanding Chemical Hazard Labels and SDS

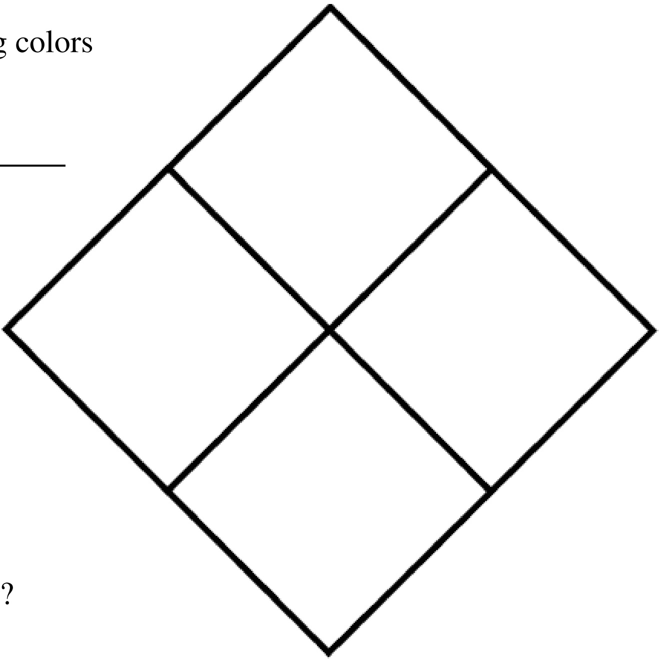
Your group will share a set of science safety papers and use it to complete this activity. DO NOT write on them, there is only one class set! When done please do not forget to place the papers back where you got them and check that the packet isn't missing any parts! Draw a star next to your name.

OBJECTIVE: To read and interpret chemical hazard labels and SDS (safety data sheets).

PROCEDURE: Use the documents found in the science safety packet to answer the following questions.

1. Interpret colors on a chemical hazard label. The following colors on a chemical label alert a user regarding what hazards?

red: _______ yellow: _______ blue: _______ white: _______



2. Interpret numbers on a chemical hazard label.

a. The number ____ is the most serious,
and the number ____ is the least serious.

b. What does the number four on a red background indicate to the user?

c. What does the number zero on a yellow background mean?

3. Use the colored pencils to fill in the appropriate National Fire Protection Agency (NFPA) hazard coding colors and numbers on the label above for potassium hydroxide. WARNING: DO NOT COPY THE NUMBERS FROM THE COLOR PICTURE PROVIDED IN THE PACKET (you will need to see section 5 of the SDS sheet to determine what the numbers should be for potassium hydroxide)

4. Use the SDS for potassium hydroxide again. These sheets must be readily available for reference by employees and laboratory students. Carefully look around the room and find where the SDS sheets are located. (They are in a large white binder with orange writing on the outside!) Describe the location of the SDS sheets in the classroom:_________________________________

Use the information given on the potassium hydroxide SDS in your packet to complete the following:

Physical appearance (i.e. what it looks like):

Melting Point:

Boiling Point:

Stability/Reactivity:

In what way is it toxic? See toxicological information.

What is the code for the Flinn suggested disposal method?

5. How to Read a Chemical Label. Find the chemical label for toluene. Use the information on the chemical label to complete the following:

First Aid:

Hazard Alert:

Shelf Life:

Part 2: Emergency equipment and situations

1. If you need to use a fire extinguisher, you will need to first pull out the metal safety pin, point the hose toward the base of the fire, and then squeeze the handle. Spray the fire extinguisher back and forth at the base of the flame until it is out. Where is the fire extinguisher?_________________________

2. If clothing catches fire do not run as this only fans the flames and makes them burn more intensely. Either put it out in the safety shower (if it's small) or stop, drop and roll if it is more serious. If your lab partner's clothing catches fire use the fire blanket to smother the fire. Where is the safety shower?____________________ Where is the fire blanket?__________________

3. Corrosive chemicals splashed on skin should be flushed off promptly with copious amounts of flowing water for at least 15 minutes. If splashed on clothing, the clothing must be removed while under the safety shower. Stay under the shower for at least 15 minutes while the instructor calls the nurse or 911.

4. A corrosive chemical in the eye is a very serious matter. Get to an eye wash fountain immediately (within 15 seconds if possible). Wash the eye for at least 15 minutes while someone else calls the nurse. Keep the eyelid open fully; move the eyeballs up, down, left, right, across, and around. How long do you use the eye wash fountain?_____________________
Where is the eye wash fountain in the room?___________________________________

5. Minimize your exposure to potential danger by wearing goggles, close-toed shoes, long pants, and tying your hair back if it's long. See #3 to remind yourself what will happen to any clothing that gets a dangerous chemical splashed on it, and re-write it here: _____________________

6. What kind of shoes should you wear in lab? ________________________ What should you do with long hair? __________________

Part 3: Lab Safety

Read the following true story that took place in a high school in California. (Kaufman, J.A.; Chemical Accidents, 1989)

In a high school in California a lab assistant mixed phosphorus and potassium chlorate ($KClO_3$) in a two-ounce gas-collecting bottle with a rubber stopper. He started to shake the bottle. The bottle exploded, blowing off one finger on his right hand. Storage bottles and containers in the storeroom where the accident occurred were pierced with flying glass. Ether was ignited by the flying phosphorus. The ether was being stored on the floor under shelving containing other organics (like gasoline), not in a safety cabinet. A student in the adjoining classroom sprayed the shirt of the assistant with a CO_2 fire extinguisher and then the assistant was taken to the nurse. The fire continued to burn for approximately fifteen minutes until the fire department arrived. The building was evacuated. All the fingers on the victim's right hand were eventually amputated. The big toe from one foot was grafted onto the thumb position so that it could press against the pad of the rest of the hand. This incident, although tragic, had potential to be catastrophic.

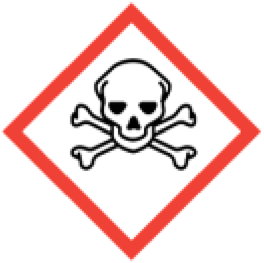
Look over the SDS sheets for potassium chlorate and phosphorus. List at least three facts that the SDS would have provided the lab assistant (if he had read them) that could have saved his fingers and/or prevented the accident. Hint: look at the sections 7, 8 and 10.

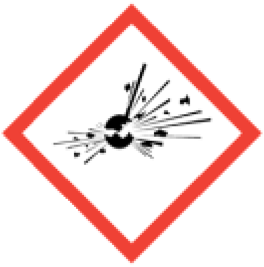
Safety symbols and additional information.
Look over the science lab safety symbols. Match each meaning to the correct safety symbol.

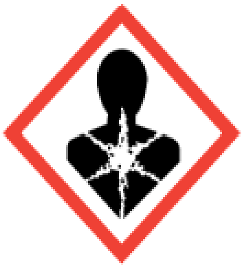
1. **_____ poisonous substance present/acute toxicity**
2. **_____ flammables, emits flammable gas**
3. **_____ corrosive chemical**
4. **_____ oxidizers**
5. **_____ carcinogen, aspiration toxicity**
6. **_____ chemicals could cause explosions**
7. **_____ gases under pressure**
8. **_____ irritant to skin/eyes (or other misc hazards)**

Read the additional information given about flammability and health hazards ("Principles of Chemical Safety") and then answer the following questions (they are not in order):

1. What is the difference between chronic and acute toxicity?


A.


B.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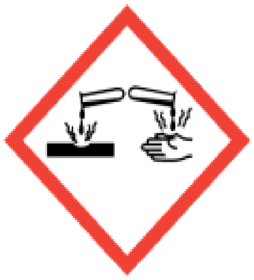
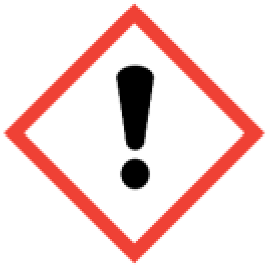

D.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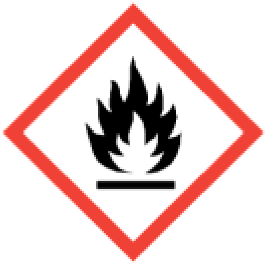
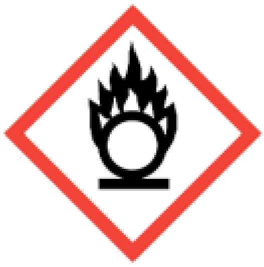
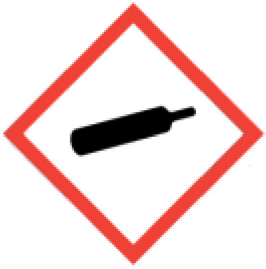

F.


G.


H.

2. Define corrosivity:

3. Which will burn more easily, gases or liquids?

4. Define volatile:

5. List the 5 principles of chemical safety (bullet points or short sentences are acceptable).

Turn in this lab and don’t forget to return all of the class sets back to the place where you got them!

Unit 2: Properties of Matter

Matter: anything that has __________ & takes up __________

Mass = measurement of the amount of _____________ an object ______________

Chemistry = the study of ______________ & the _____________ that it undergoes

The Three States of Matter

Solid: _______________ shape (independent of the container), definite volume (cannot be ____________), particles are _____________________________ and cannot move

Liquid: _________________ shape (depends on container), _____________ volume, particles packed together but ___________________

Gas: indefinite volume (can be ___________________), indefinite ____________, particles ____________________________ & move _______________________

Energy: ability to _______________ OR cause a _____________ in ________________

Work = moving _____________ from one ______________ to ______________

Potential Energy = _________________________. This is what makes _______________ reactions possible

Mixtures: a combination of _____ or more types of _________________

Each component keeps its own ______________ & ___________________

Components are only ___________________ mixed & can be _________________ using _______________ means

Types of Mixtures

1. ________________________ mixture -

a ________________ in which the substances are _______ evenly distributed

2. ________________________ mixture –

a type of mixture (AKA '_________________) in which the substances are __________________________________

Separating Mixtures

Exploits differences in _____________________________

1. ___________________ – a technique that uses a porous ________________ to separate a _____________ from a ______________

2. ___________________ – a technique that can be used to _________________ separate most ______________________ mixtures based on differences in the ____________________ of the substances involved

3. ______________________ – a technique that is used to _________________ separate the ____________________ of a mixture based on the tendency of each component to travel or be drawn across the surface of another material

Pure Substances

Every sample has same characteristic _________________ and _________________

If made of:

_____ type of atom: ___________________ (Ex: iron, gold, oxygen)

____ or more types of atoms: _________________________ (Ex: salt, sugar, water)

Element: simplest form of ________ that has its own unique set of ___________.

Cannot be separated into simpler substances by __________________ or __________________ means

Each element is made up of ____

- can be _________________ elements (e.g. H_2, S_8, O_3) or ____________ elements (e.g. Al, Au, Fe)

Compound: a substance that contains ____ or more ______________ chemically combined in a _______________________________

Properties are ______________ from the ______________________________ ex. Glucose (sugar)

Breaking Down Compounds

__________________ methods do not work

Only __________________ changes can ___________ compounds down into other types of ______________

I. Classifying Matter

A. Quantitative vs Qualitative

1. Quantitative – ________________ information (e.g. how much, how little, how big, how small)

2. Qualitative – ________________ information (condition, ____________, size, ___________, odor, texture)

B. Physical Properties – a characteristic that can be observed or measured without changing the _________________ of a substance

ex. ___________, _____________, taste, hardness, _____________, melting/boiling points, ______________________

II. Chemical Changes vs Physical Changes

A. States of Matter – the physical __________ in which all matter exists on Earth

B. Physical Changes – a change that occurs in the ______________________________________ of a substance, but does not change its chemical____________

Ex. ______________, __________________, _________________, evaporating, _________________, condensing, breaking, splitting, grinding, cutting, ______________________

C. Chemical Properties – relates to a substance's ________________ to undergo changes that transform it into a __________________ substance

- easiest to see when a chemical is ________________

D. Chemical changes – a change that produces matter with a different _______________________________ than the original matter

ex. ____________, rotting, ______________, reacting, cooking, _________________, respiration

E. Chemical Reactions

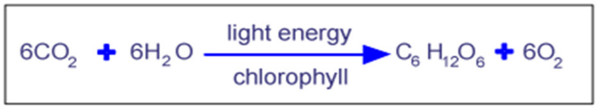
The process by which the ___________ of one or more substances are _________________ to form _______________ substances

- Also known as a ________________________________
- Chemical properties can only be ________________ when a substance undergoes a ________________________

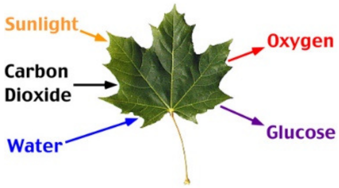
Example of chemical change:

The process of __________________________ is a chemical reaction in which light energy, carbon dioxide and water, are transformed to create glucose and oxygen. The chemical reaction can be shown by writing out the element symbols for each compound.

Instructions: label the products and reactants for the photosynthesis reaction below.



$$6CO_2 + 6H_2O \xrightarrow[\text{chlorophyll}]{\text{light energy}} C_6H_{12}O_6 + 6O_2$$



F. Recognizing Chemical Changes

1. How can you tell whether a chemical change has taken place?

a. Transfer of __________________ (Energy may be given off in the form of ________ or __________)

b. Change in _________ (changes in molecular ________________ may cause substances to reflect ___________ differently)

c. Precipitate (A ________________ that settles out of a ____________ mixture)

d. __________ production (indicated by the formation of ______________________)

III. Law of Conservation of Mass

A. In any physical or chemical change, ____________ is conserved

i.e. Mass is neither __________________ nor ____________________

B. Examples of conservation of matter:

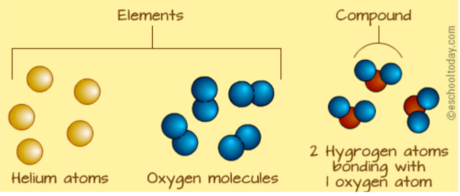
1. In chemical reactions, the mass of the _________________ is always equal to the mass of the _____________________

2. The total _____________ does not change in ____________ changes, such as the _______________ of ________

I. Nature of Materials

A. Most natural materials are _________________ of _________________________

B. Pure substances can be either _________________ or ___________________



1. Element

a. A substance composed of ____

b. All elements above _______ are artificial

c. Examples: ___________________________________

2. Compound

a. A substance composed of ____

b. Examples: ___________________________________

II. Names and Symbols for the Elements

A. Each element has a name and a symbol

1. Symbol is _____ or _______ letters

2. First letter always ____________________

3. Second letter always ___________________

B. Symbol usually consists of the first one or two letters of the element's name.

Examples - Oxygen: ______ Krypton: ______

C. Subscripts indicate the relative _______________________ of the _______________ in a compound

D. Sometimes the symbols come from original Latin or Greek names

1. Examples: gold: ______ = aurum (shining down) lead: ______ = plumbum (heavy)

E. Sometimes names reflect the place where it was discovered or to honor a famous scientist

Examples: Einsteinium: ____ (Albert Einstein) Californium: ____ (University of California at Berkeley)

III. Dalton's Atomic Theory

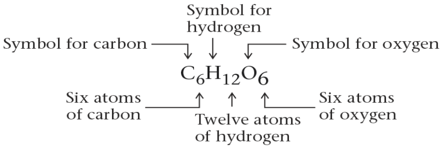
A. All elements are composed of ________________

B. All ______ of a given element are _____________________ (he didn't know about ________________).

C. _____________ of different elements are _____________________

D. Compounds consist of the ______________ of different _________________

E. Atoms are neither __________________ nor ________________ in a chemical reaction.



IV. Chemical Formulas

A. Represent the ____________ and ________________ of atoms present Ex: ________________, a simple _______________________

B. Rules for writing formulas

1. Each _____________ present is represented by its element symbol.

2. The ________________ of each type of atom in the molecule is indicated by a ____________________ written to the ___________ of the element symbol.

3. When only one atom of a given type is present, the subscript is _____ written.

C. Practice writing chemical formulas: write the correctly-formatted formulas below:

1. A molecule contains four phosphorus atoms and ten oxygen atoms

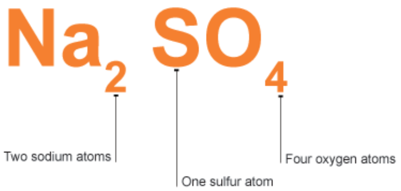
2. A molecule contains one uranium atom and six fluorine atoms

3. A compound contains one aluminum atom for every three chlorine atoms

4. A pure element consists of helium atoms not bonded together

5. A pure element consists of two fluorine atoms bonded together

Instructions: label the subscript(s) in the formula below. This means write the word "subscript" and draw arrow(s) pointing towards the appropriate things



Name: __________________ Per: ____

Note Questions: Unit 2 - Properties of Matter

Notes #1 - Matter

1. What are the two characteristics of matter?
    a.
    b.

2. What is mass a measurement of?

3. What are the three states of matter? ____________ and ___________ and _____________

    a. Which state of matter has definite shape and definite volume?

    b. Which state of matter has indefinite shape and definite volume?

    c. Which state of matter has indefinite shape and indefinite volume?

4. What two things can energy do?
    a.
    b.

5. What is a heterogeneous mixture?

6. What is a homogeneous mixture?

7. Which way of separating mixtures is based on using different boiling points of substances in a mixture to separate them?

8. Which way of separating mixtures is based on the tendency of each component to travel or be drawn across the surface of another material?

9. Which type of separation would work best to separate the salt out of the water in salt water?

10. Define "element" in terms of the kind(s) of atoms it contains

11. In general, can elements be broken down into simpler substances?

12. Define "compound" in terms of the kind(s) of atoms it contains.

13. In general, can compounds be broken down into simpler substances?

Notes #2 – Properties of Matter

14. What are "physical changes" in a substance?

15. Give at least four examples of physical changes
    a.
    b.
    c.
    d.

Be sure to complete all questions on the next page as well!

16. What are “chemical changes” in a substance?

17. Give at least four examples of chemical changes
    a.
    b.
    c.
    d.

18. What is a chemical reaction?

19. What are four potential signs that a chemical reaction has taken place?
    a.
    b.
    c.
    d.

20. The Law of Conservation of Mass says what about matter?

21. Is the Law of Conservation of Mass true for chemical, physical changes, or both?

Notes #3: Elements, Compounds and Chemical Formulas

22. Are most substances mixtures or pure substances?

23. What is the difference between an element and a compound?

24. Is the first letter in the symbol for an element sometimes, always, or never capitalized? __________
    What about the second letter in the symbol, is it sometimes, always, or never capitalized? __________

25. Circle the subscript(s) in this chemical formula:

$$Na_2HPO_4$$

26. Does a chemical formula represent a total number or a ratio of atoms?

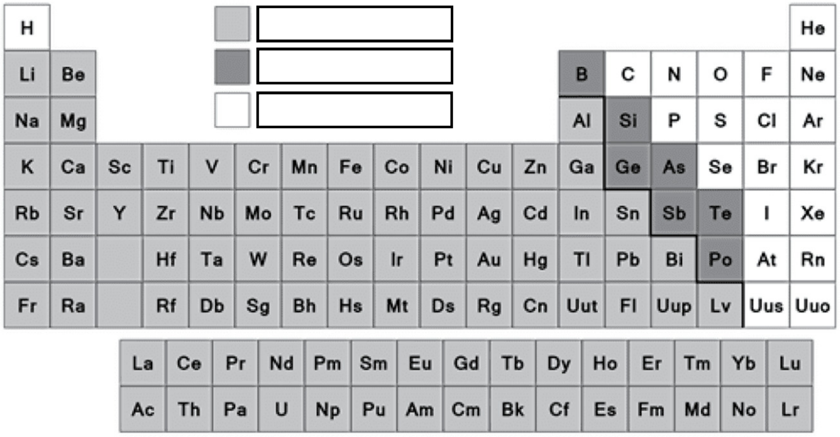
Unit 3: Atomic Structure and the Periodic Table

Name: ______________________ Per: ____

Elements of the Periodic Table

In the boxes for the color key below, label which elements on the periodic table below are metals, non-metals, and metalloids. If given a blank, uncolored, unlabeled periodic table having only the element symbols on it, you should be able identify whether a given element is metal, non-metal or metalloid.

Fill in these labels!

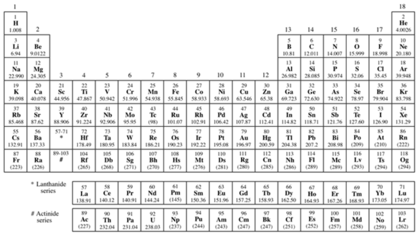

H
He
Li Be
B C N O F Ne
Na Mg
Al Si P S Cl Ar
K Ca Sc Ti V Cr Mn Fe Co Ni Cu Zn Ga Ge As Se Br Kr
Rb Sr Y Zr Nb Mo Tc Ru Rh Pd Ag Cd In Sn Sb Te I Xe
Cs Ba Hf Ta W Re Os Ir Pt Au Hg Tl Pb Bi Po At Rn
Fr Ra Rf Db Sg Bh Hs Mt Ds Rg Cn Uut Fl Uup Lv Uus Uuo
La Ce Pr Nd Pm Sm Eu Gd Tb Dy Ho Er Tm Yb Lu
Ac Th Pa U Np Pu Am Cm Bk Cf Es Fm Md No Lr

Below is a list of the elements we will commonly encounter in this class. Because we use them frequently, you need to memorize their names and symbols. It is not necessary to memorize the element numbers, but do include them on this worksheet for ease of reference. Some have been done for you as an example, please follow that pattern. For elements with differing British/American spellings, please use the American spelling (Aluminum and Sulfur have been spelled for you due to this concern).

Element Symbol	Element Name	Element Number	Metal, Non-metal or Metalloid?		Element Symbol	Element Name	Element Number	Metal, Non-metal or Metalloid?
Ag	Silver	47	Metal		Li			
Al	Aluminum	13	Metal		Mg			
Ar					Mn			
As					N			
Au					Na			
B					Ni			
Ba					O			
Be					P			
Br					Pb			
C					S	Sulfur		
Ca					Si			
Cl					Sn			
Co					Sr			
Cr					Zn			
Cu								
F								
Fe								
H								
He								
Hg								
I								
K								

Name: ______________________________ Per: _____

Explainer: What is an Isotope?



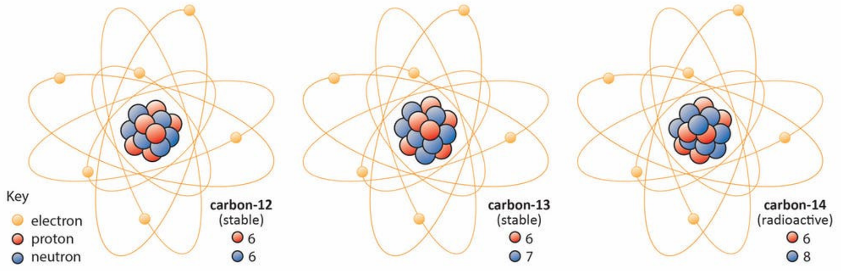
If you've ever studied a periodic table of the elements, you're probably already aware that this table reveals a great deal about the chemical properties of the atoms that make up our world. But you may not realize that each square on the periodic table actually represents a family of isotopes — versions of the same element which share the same chemical identity and properties, yet have different masses. To better understand what isotopes are, we need to take a closer look at the interior of an atom.

Building blocks of matter

In order to understand isotopes, we need to understand what atoms are. An atom is composed of an incredibly dense core (called a nucleus) of protons and neutrons, surrounded by a diffuse cloud of electrons. You can think of protons and neutrons as the same kind of particle with one key difference: the protons are positively charged, while neutrons carry no charge. They are held together by the nuclear strong force. The electrons, which are much lighter than protons or neutrons, carry the same magnitude of charge as a proton but with the opposite sign, meaning that each atom that has equal numbers of protons and electrons is electrically neutral. The attractive forces between these opposite charges keep the electrons from leaving the atom.

Atoms and isotopes

Isotopes are forms of an element that have the same number of protons but different numbers of neutrons. This is because the identity of an element is determined by the number of protons it has. Let's use element #6 (carbon) as an example: one form of carbon has a mass number of 12, which means it has 6 protons and 6 neutrons (note that the element number, AKA the atomic number, tells you how many protons there are). Another form has a mass number of 13, which means it has 6 protons and 7 neutrons, and yet another form has a mass number of 14, which means it has 6 protons and 8 neutrons.



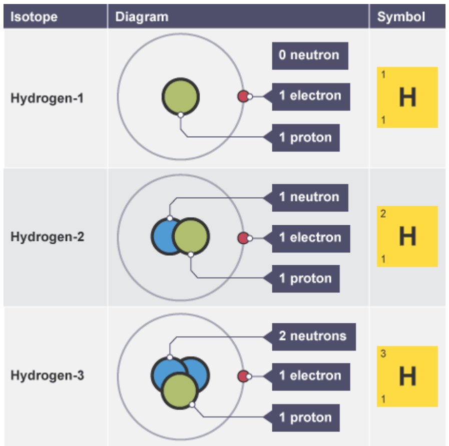
How to Denote Isotopes

The word "isotope" comes from the Greek words *isos* (equal) and *topos* (place), which signify that isotopes of an element occupy the same place in the periodic table, even though they have different atomic masses. Unlike atomic number, which represents the number of protons in the nucleus, atomic mass is the mass of all the protons and the neutrons combined.

One way to denote an isotope is to write the symbol of the element followed by a number that denotes the total number of particles in its nucleus. For example, the isotope of carbon (atomic #6) with 6 protons and 6 neutrons in its nucleus is written as Carbon-12. Another isotope of with 6 protons and 8 neutrons would be written as Carbon-14 since it has 14 particles in the nucleus.

Another way to denote isotopes is with subscripts and superscripts before the symbol of the element, where the subscript is the number of protons and the superscript is the atomic mass. Using this method, you would denote carbon-12 as $^{12}_{6}\mathrm{C}$ and carbon-14 as $^{14}_{6}\mathrm{C}$. Nitrogen-14 (it's atomic # is 7, so it has 7 protons) would be $^{14}_{7}\mathrm{N}$.

This same pattern applies to all elements, including the three natural isotopes of hydrogen (atomic #1), as seen in the diagram at right.

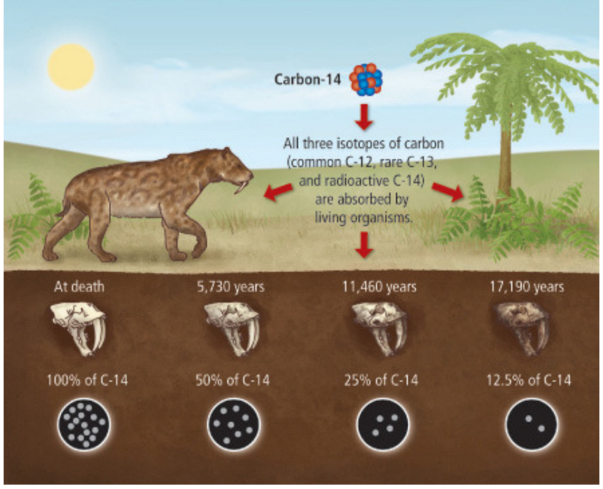


Isotope	Diagram	Symbol
Hydrogen-1	0 neutron 1 electron 1 proton	$^{1}_{1}\mathrm{H}$
Hydrogen-2	1 neutron 1 electron 1 proton	$^{2}_{1}\mathrm{H}$
Hydrogen-3	2 neutrons 1 electron 1 proton	$^{3}_{1}\mathrm{H}$

Average Atomic Mass

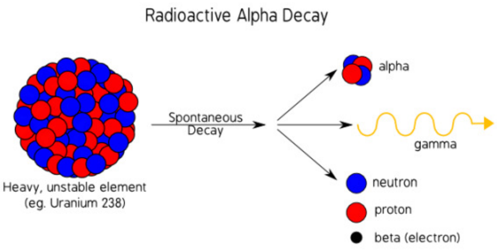
Every element that occurs in nature has multiple isotope forms, and scientists have managed to synthesize many more in the laboratory. In all, there are 275 isotopes of the stable elements and around 800 radioactive isotopes. Because each isotope has a different atomic mass, the atomic mass listed for each element in the periodic table is an average of the masses of all the isotopes weighted by the total percentage of each isotope that occurs in nature.

For example, in its most basic form, the hydrogen nucleus consists of just a single proton, and is called protium $({}^{1}_{1}\mathrm{H})$. There are, however, two other naturally occurring isotopes: deuterium $({}^{2}_{1}\mathrm{H})$, which has one neuton, and tritium $({}^{3}_{1}\mathrm{H})$, which has two. Because the form containing no neutrons is by far the most abundant, the average atomic mass of all hydrogen is close to that of protium, with a value of 1.008 AMU.


Carbon-14
All three isotopes of carbon (common C-12, rare C-13, and radioactive C-14) are absorbed by living organisms.
At death
5,730 years
11,460 years
17,190 years
100% of C-14
50% of C-14
25% of C-14
12.5% of C-14

Marching towards stability

Every element is known to have isotopic forms with identical chemical properties. Some isotopes however, have an imbalance where having too many or too few neutrons can cause the nucleus to become unstable. If the nuclear strong force isn't strong enough to hold protons and neutrons together, the atom will eject neutrons or other energetic particles, and with them, a certain amount of energy. This process is called radioactivity.


Radioactive Alpha Decay

Due to alpha decay, two protons and two neutrons are emitted. In the case of Uranium 238, it decays into Thorium-234.

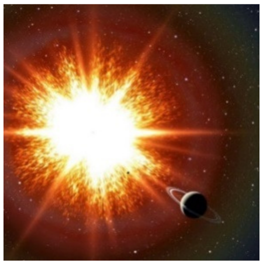
It is through the process of radioactivity that unstable isotopes decay until they reach stability. They do this until they are no longer unstable, sometimes going through multiple steps and transforming into different isotopes, or even different elements along the way. Stable isotopes, by contrast, show no tendency to undergo radioactive decomposition and will remain perpetually unchanged under natural circumstances. One example is carbon-12, which will happily continue to exist as carbon-12 unless something extraordinary happens. Carbon-14, which is radioactive, will at some point decay into a stable isotope by changing one of its neutrons into a proton, forming nitrogen-14. During this process, which is known as beta decay, the nucleus emits radiation in the form of high-energy electrons.

We have found ways to make use of naturally occurring radioactive isotopes to find out how old an object is. Carbon dating, for example, makes use of the carbon-14 isotope to determine how old biological objects are.

Under normal circumstances, carbon-14 is produced in our atmosphere via cosmic ray alterations of nitrogen-14. It has a half-life of roughly 5,700 years, which means that half of a quantity of carbon-14 will have decayed away in that time period. While a biological organism is alive, it takes in approximately one carbon-14 isotope for every trillion stable carbon-12 isotopes, so the ratio of carbon-14 to carbon-12 stays about the same while the organism lives. Once it dies, new intake of carbon stops, and as the carbon-14 decays away, the ratio of carbon-14 to carbon-12 decreases in the remains of this organism over time. If we extract the carbon from the sample using chemical methods, we can then use accelerator mass spectrometry (AMS) to find the ratio of carbon-14 to carbon-12, and based on that, estimate the amount of time since the sample was last part of a living organism.

Origins

By now, you may be wondering how all these isotopes were created in the first place. As it turns out, this question is a complex one, but lends some truth to the adage that we are all made of star dust. Although some isotopes are formed by the collisions between cosmic rays and the gases in our atmosphere, most trace their origin back billions of years. The lightest atoms originated 13.7 billion years ago in the Big Bang, while other, heavier ones resulted from processes of fusion that happen within stars. Our sun fuses hydrogen into helium, but other, larger stars are hot enough to fuse hydrogen and helium into much heavier elements up through element 26, which is iron. When those larger stars collapse and explode at the end of their lives, all of the heavier elements - including radioactive isotopes – are formed in a gigantic explosive supernova which sends them outward at the incredible speed of over 20,000 kilometers per second (by comparison, a supersonic jet fighter can reach peak speeds of only around 0.5 km/sec). This process seeds the universe with all of the natural elements up through uranium and is the source of all the dust that formed our solar system, our planet, and everything on it. That means no matter what atom you look at – even the ones in your own body – each is incredibly old and dates back billions of years to a time before there even was a planet earth. Thus, even if you are young, your building blocks are incredibly old!



Name: __________________________ Per: ____

Dialectical Journal for “Explainer: What is an Isotope?”

<table>
<tr><th>Quote/Passage/Summary/Paraphrase</th><th>Connection</th></tr>
<tr><td><u>Paragraph 1 - Quote 1</u>
“Each square on the periodic table actually represents a family of isotopes”</td><td>Read the paragraph, and then explain in more detail what is meant by the quote:
____
____
____
____
____
____
____
____</td></tr>
<tr><td>Paragraph 2</td><td>Explain the structure and composition of an atom by describing the characteristics of each of the following:<table><tr><th><u>Protons</u></th><th><u>Neutrons</u></th><th><u>Electrons</u></th></tr><tr><td></td><td></td><td></td></tr><tr><td></td><td></td><td></td></tr><tr><td></td><td></td><td></td></tr><tr><td></td><td></td><td></td></tr><tr><td></td><td></td><td></td></tr><tr><td></td><td></td><td></td></tr></table></td></tr>
<tr><td>Paragraph 3</td><td>Compare and contrast different isotopes of the element carbon. In other words, explain what they have in common, and what differs between them.
____
____
____
____
____
____
____
____</td></tr>
<tr><td>Paragraphs 4 - 6

From paragraph 6: “This same pattern applies to all elements, including the three natural isotopes of hydrogen, as seen in the diagram…”</td><td>Show that you understand the quote at left by <i>drawing</i> two lithium isotopes (Li = atomic #3), one with a mass of 7 AMU and one with a mass of 6 AMU. For each picture, label the neutrons, protons and electrons and give the name and symbol for the isotope.<table><tr><td>Lithium isotope with a mass of 6
Isotope name: ________________
Isotope symbol: _______________
Diagram with labels:</td><td>Lithium isotope with a mass of 7
Isotope name: ________________
Isotope symbol: _______________
Diagram with labels:</td></tr></table></td></tr>
</table>

Paragraphs 7-8 Use oxygen’s most stable isotopes for the question: Oxygen - 16 99.76% Oxygen - 17 0.03037% Oxygen - 18 0.204%		The periodic table lists a mass of 15.999 AMU for oxygen. Explain why it isn’t closer to 17 or 18. Use academic vocabulary from paragraph 8.
Paragraphs 9 - 10		Explain the distinction between stable and unstable isotopes:
Paragraphs 11-12 “…we can … find the ratio of carbon-14 to carbon-12, and based on that, estimate the amount of time since the sample was last part of a living organism.”		Explain why the ratio of carbon-14 to carbon-12 tells you how old something is:
Paragraph 13 “…but lends some truth to the adage that we are all made of star dust”		Explain the meaning of this quote, citing evidence from the reading

Summarize your knowledge about isotopes:

Name: _________________________ Per: ____

Understanding the Cathode Ray Experiment

Part 1: Understanding the experiment

Open the video "Discovery of the Electron Cathode Ray Tube Experiment"

1. What question was JJ Thompson trying to answer?

2. A cathode ray tube is a tube of ___________ which is sealed all over, and the ______________ has been pumped out of it.

3. Sketch and label the cathode ray tube (switched on with cathode ray showing) as it appears in the video.

Now open the video "RC Unit 3 Demo - Cathode Ray Tube" so that you can see what a real cathode ray tube looks like.

4. Describe what happens to the cathode ray as it is exposed to magnetic or electrical forces:

Now return to and resume watching the video "Discovery of the Electron Cathode Ray Tube Experiment"

5. What did Thompson put on the outside of the cathode ray tube that caused the cathode ray to bend?

6. Which charge did the cathode ray bend toward?

7. Based on your answers to questions 5 and 6, what could be concluded about the charge of the particles in a cathode ray? _________ Now explain the reasoning behind this answer in detail (no credit for a short answer here since that probably won't be long enough to contain sufficient detail)

8. Thompson decided to check his results another way. What did he use?

9. During this second experiment, the cathode ray bent a certain way. As a result, Thompson was able to conclude what about the charge of the particles in the cathode ray?

10. The video mentions “other experiments” that Thompson did. This includes calculating the mass-to-charge ratio, from which he was able to deduce the mass of the particles in the cathode ray. What did he find out about their mass?

11. Regardless of what metal he used, what did Thompson notice about the particles in the cathode ray?

12. Based on his three observations, what was Thompson able to conclude?

*Hint: “atoms have ____

____

The particles are called ____________________________

13. Explain How Thompson knew that atoms contain some positive charge:

14. Thompson’s “plum pudding” model of the atom
a. Sketch and label the “plum pudding” model here:

b. Describe the “plum pudding” model here:

Part 2: Interacting with the cathode ray tube experiment

Reach the online simulation:

a. Go to the following URL: https://learn.concord.org/resources/759/crookes-tube
b. Click the orange "Preview" button

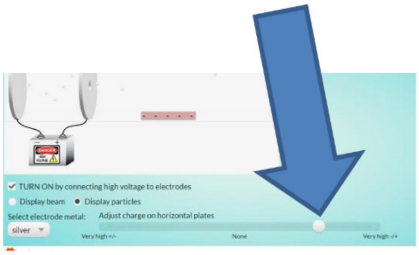
1. Click the "TURN ON" button. You should see a green beam appear. Now click "display particles". Place a check mark here ____ when you have done this.

2. According to what you learned in the video earlier, what are these particles?

3. Why are they being deflected upwards initially?

4. In the lower right is a slider (see picture). Move the slider around. What is it adjusting in the cathode ray tube?

5. How do the particles react when you adjust the slider?



☑ TURN ON by connecting high voltage to electrodes

○ Display beam ● Display particles

Select electrode metal: silver

6. The particles are always attracted to __________ charge, therefore, they must have a charge that is ___________

(word bank: positive, negative)

7. Now change the type of metal in the lower left part of the screen. Do the particles look or behave differently? These metals are the source of the particles you see flying across the screen. Therefore, all metals that Thompson tested contain particles that are same as each other or different? _________________

8. Now click "display beam". The particles are too small to see in real life, so working in the lab, you would see the green beam. Move the slider back and forth, noticing how it behaves in the exact same way as the particles since it is composed of them. Place a check mark here ____ when you have done this.

9. Write a summary explaining how J.J. Thompson was able to experimentally demonstrate that all of the atoms he tested contain electrons.

Name: ______________________________ Per: ______

The Gold Foil Experiment

Understanding how we came to know that atoms have a nucleus

1. Start by reading this introduction:
The atom is far too small to see, and even to this day it is impossible to directly see the nucleus or anything in it. So how do we know that atoms are made of, or even that they contain a nucleus? The answer comes from carefully-designed experiments that allowed scientists to infer that these things must exist due to the results they observed. It took many tries and mistakes, and sometimes the models weren't completely right, but their models were based on their evidence, and each model was more accurate than the one before thanks to the data provided by these experiments.

a. Although we can't directly see the nucleus of an atom, how do we know that it exists?

b. How do scientists increase the accuracy of their models?

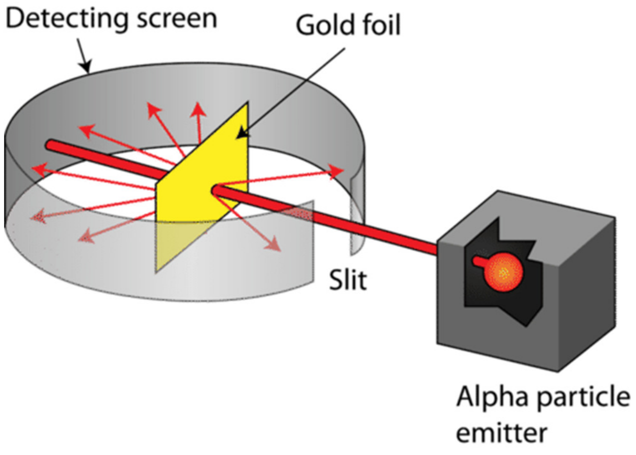
Now start the video "Discovery of the Nucleus: Rutherford's Gold Foil Experiment" (You will answer up through all of question 3 before pausing at 4:55)

2. After JJ Thompson conducted his cathode ray experiment, he used his data observations to create his "plum pudding" model of what he thought the atom's structure was. Draw this model, take a picture of your drawing, and insert the picture below:

3. Rutherford set out to test whether or not the "plum pudding" model of the atom's structure was correct. He set up an experiment in which he would shoot alpha particles at the gold foil and measure how they interact with it.

a. Define what an alpha particle is AND specify what electrical charge it has:

b. At right is a diagram of what the experimental set-up looked like. If you were to draw an arrow showing which way the alpha particles are moving, would they be moving toward or away from the alpha particle emitter? _______. Place a capital X in this blank here _____ once you have read this instruction.



c. Where did most of the alpha particles go?

d. A few of the alpha particles didn't go in the same direction as the rest. What happened to them?

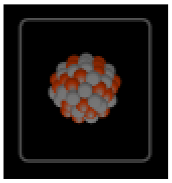
Now pause the video at 4:55. You will resume it later.

4. Before doing the gold foil experiment, Rutherford had certain expectations about what he thought he was going to see. Let's look at a simulation to help you see it too.

a. go to this web address:
https://phet.colorado.edu/sims/html/rutherford-scattering/latest/rutherford-scattering_en.html

b. Click (or double-click) "plum pudding atom"
c. Click the blue "alpha particles" button on the left
d. Now you should see positively-charged alpha particles moving across the screen in an upward direction.
e. Are the alpha particles moving in a straight line, or are some of them changing direction as they move through the atom? ____________________________________
f. Is the positive charge in the atom collected together in one spot, or is it spread throughout the atom?
____
g. Alpha particles are positive, and so is part of the atom. Do like charges attract or repel?
____
h. Based on what you're seeing, is the positive charge in an atom concentrated enough to push any of the bullet-like alpha particles off course? _____
Does this simulation match what Rutherford saw in the actual experiment? _________

5. Now let's look at a simulation showing what happened in the actual Rutherford atom experiment.
a. click the "Rutherford atom" icon at the bottom of the screen
b. Click the blue "alpha particles" button on the left
c. The alpha particles are purple. Are some of them moving in a straight line from bottom to top? _______ Are some of them not moving in a straight line, as if they were bouncing off of something? _________
d. Now click the nucleus picture in the upper left (looks like the picture at right



e. alpha particles have a positive or negative charge? ___________
f. Describe what the alpha particles that get close to the nucleus are doing:
____
g. In order to explain your observations, the charge of the nucleus must be positive or negative?
____
h. What would the alpha particles do if they had the opposite charge as the nucleus?
____
i. On the right side of the screen, reduce the number of protons in the nucleus to 20.
j. Describe how the behavior of the alpha particles changes: ____________
k. Return the number of protons back up to 79. Watch the alpha particles flow for a bit. Then reduce the number of neutrons to 20. Place a check mark here when you have done this: _______

l. You should have observed that reducing the number of neutrons from 118 to 20 didn’t change the behavior of the alpha particles. Explain why reducing the number of protons changed their behavior, but reducing the number of neutrons did not:

Now resume the video “Discovery of the Nucleus: Rutherford's Gold Foil Experiment” from 4:55

6. What part of the experimental results showed that most of the atom is composed of empty space?

7. What part of the experimental results showed that the center of the atom is small and positively-charged?

8. Draw and label the model of the atom proposed by Rutherford to explain his results (see around 9:10 in the video).

9. Write a short paragraph of a few sentences describing the model proposed by Rutherford.

10. Watch Mr DeWitt's summary of how the model of the structure of the atom changed over the years.
    a. Was Rutherford aware that the nucleus's positive charge was from multiple, individual protons? ____
    b. When were neutrons discovered? _________

11. Explain why some of the alpha particles bounced back instead of going straight through during the real-life experiment.

12. Dalton (around the year 1810) thought that atoms were completely solid masses. Consider the observations with real-life atoms in comparison with what would theoretically happen if atoms had no nucleus at all. What would the alpha particles have done if the Dalton model of atoms being totally solid masses was true? The video doesn't say, but you should be able to come up with the answer yourself.

    a. Draw a diagram

    b. Write your answer here:

I. Subatomic particles

A. ______________________________

1. Located in ______________________

2. All atoms of same _________________ have the same _______________ of ________________

3. _____________________________ = number of protons

B. _______________________________

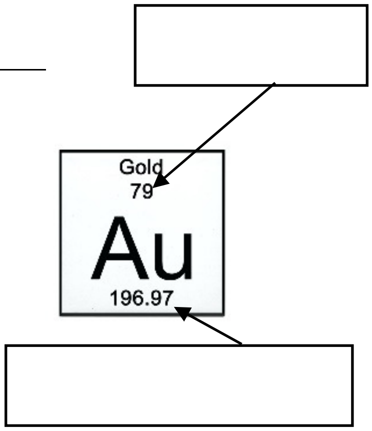
1. Surround the __________________

2. Number of electrons = number of ___________ if atom is ________________

C. _______________________________

1. # of neutrons _____________________

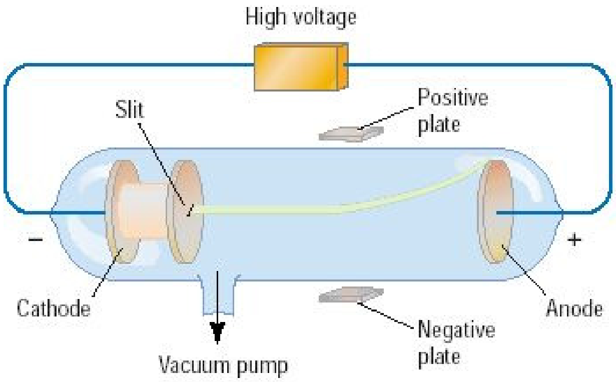
2. Mass number = sum of __________________ and ________________



D. Change to Dalton's atomic theory

1. All atoms of the same element contain the same # of _________________ and _______________, but can have a different # of __________________

II. Discovery of the _________________________



A. In 1897, J.J. Thomson used a cathode ray tube to detect the presence of _________________charged ____________________

1. The stream of _________________ was _______________ when charged metal ____________________ were placed outside the _________________

2. How Thompson knew that __________________ are _________________: they veered toward the _____________ plate and away from the ___________________ plate

3. Same _________ to ___________ ratio for all ___________ tested, so they must all ___________ this ______________

4. Conclusions from the Study of the Electron

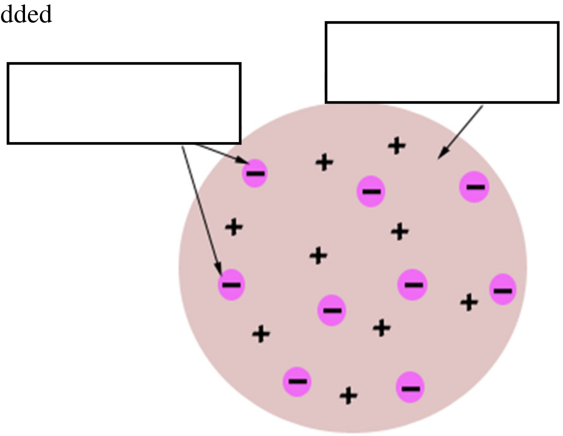
a. Cathode rays have identical ___________________ regardless of the ____________________ used to produce them. All elements must contain ______________________________________

b. Atoms are ________________, so there must be __________________ particles in the atom to balance the _______________ charge of the electrons

c. Electrons have so little ___________ that atoms must contain other _______________ that account for most of the _________

B. Thomson's Atomic Model

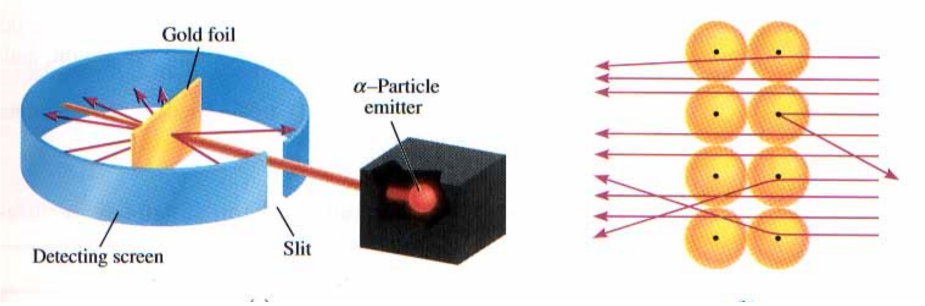
1. Thomson believed that the electrons were like _____________ embedded in a _______________ charged "pudding," so it was called the "______________________________" model.



III. Rutherford's Gold Foil Experiment

A. Alpha particles (__________________________) were fired at a thin sheet of __________________________

1. Particle hits on the detecting screen (film) are ______________________



B. Rutherford's Findings

1. _____________ of the particles ______________________________________

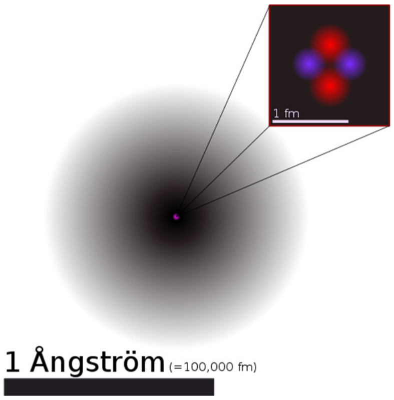
2. A few particles were _____________________

3. Conclusions:

a. The nucleus is _____________________

b. The nucleus is _____________________

c. The nucleus is ________________________________



IV. Subatomic Particles

Particle	Charge	Mass (kg)	Location
		9.109×10^{-31}	
		1.673×10^{-27}	
		1.675×10^{-27}	

A. The Atomic Scale

1. Most of the mass of the atom is in the ________________________ (__________________ and _____________________)

2. ____________________ are found ______________________ of the nucleus (the electron ___________________)

3. Most of the volume of the atom is ________________________________

B. The identity of an element is based on how many __________________ it has

Example: _______________ always has ____ protons

1. The number of __________________ in an element can _____________

Example: some forms of carbon have _____ neutrons, and other might have ____, ____ or _____

V. Isotopes

A. Atoms with the same number of ____________________ but different numbers of _____________________

B. Can be represented in two ways:

a. symbol with mass and atomic numbers

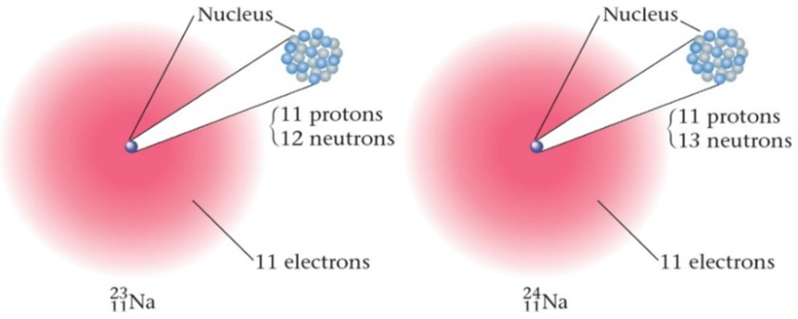
b. Element name – mass number

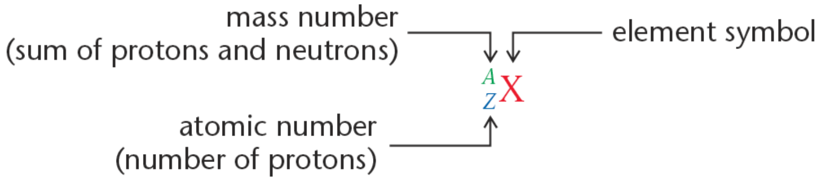
Example: Sodium – 23 (Symbol = $^{23}_{11}$Na)

# of Protons: atomic number = ____

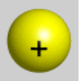
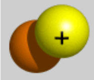
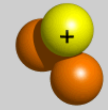
# of Neutrons = Mass # - atomic # = __

# of Electrons = protons = ______





Isotope example: different forms of hydrogen

Isotope Name	Protons	Electrons	Neutrons	Nucleus	Isotope Symbol
Hydrogen–1 (protium)					
Hydrogen-2 (deuterium)					
Hydrogen-3 (tritium)					

VI. Atomic Masses

A. Atomic mass is the average of all the natural isotopes of that element.

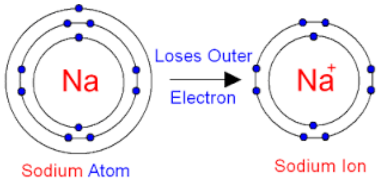
Isotope	Symbol	Composition of the nucleus	Percent in nature
Carbon-12			98.89%
Carbon-13			1.11%
Carbon-14			<0.01%

Overall average for carbon: ______________________

I. Ions

A. Atoms can form ions by ________________ or __________________ electrons

1. Ions are NEVER formed by changing the number of ___________________



B. Cations

1. ___________________________

2. ____________ tend to ____________ one or more ________________

3. Generally named by using the __________ of the ___________________

Ex: ____

Ex: ____

Ex: ____

Cation
cat•i•on
Pronunciation: [kat-ahy-uh n, -on]
-noun, *Chemistry*
1. An ion with a paws-itive charge.
2. The cutest ion ever.

Cation … the cutest ion ever!

C. Anions

1. ____

2. ________________ tend to _________ one or more ________________

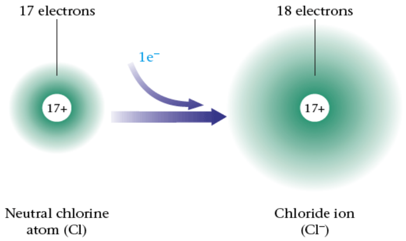
3. Anions are named by using the root of the _____________ name

followed by the suffix " ____________"

Ex: _________________________

Ex: _________________________

Ex: _________________________



D. Practice calculating numbers of protons, neutrons and electrons for ions. Show work for calculations (just like the examples) and complete the entire table

Chemical Symbol	Number of protons	Cation or anion?	Electrons gained or lost?	Number of electrons (show your work)	Number of neutrons (show your work)
K^+	19	Cation	Electrons lost	19 - 1 = 18 electrons	39 - 19 = 20 neutrons
$^{127}_{53}I^-$					
Mg^{2+}					
$^{15}_{7}N^{3-}$					
$^{28}_{13}Al^{3+}$					
O^{2-}					

II. Ion Charges and the Periodic Table

A. Ion _______________ can often be predicted from the periodic table.

Reminder: Groups are the numbers at the top of each column of the periodic table

Group 1 forms _____ ions

Group 2 forms _____ ions

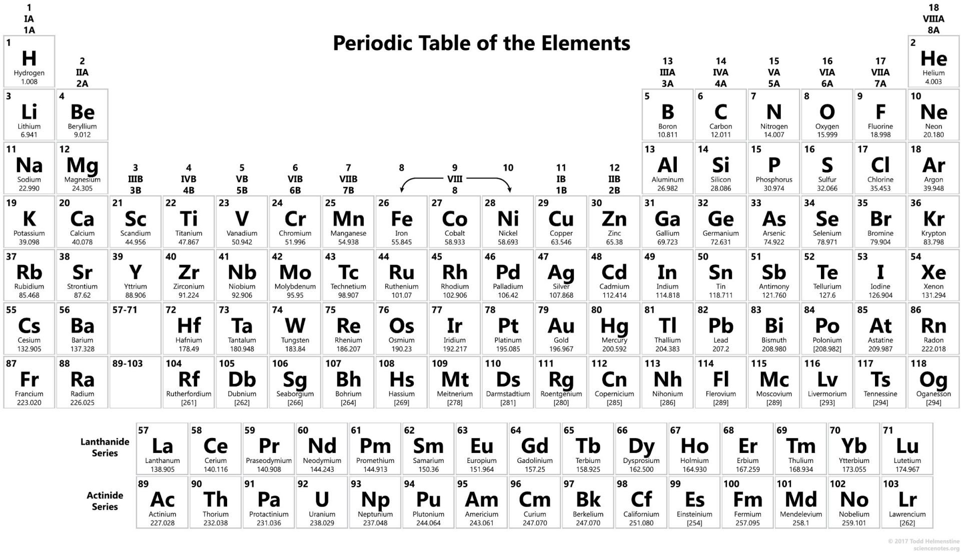
Group 13 forms _____ ions

Group 15 forms _____ ions

Group 16 forms _____ ions

Group 17 forms _____ ions

Label the charges for each column as directed by your instructor


Periodic Table of the Elements
1 IA 1A
2 IIA 2A
3 IIIB 3B
4 IVB 4B
5 VB 5B
6 VIB 6B
7 VIIB 7B
8
9 VIII 8
10
11 IB 1B
12 IIB 2B
13 IIIA 3A
14 IVA 4A
15 VA 5A
16 VIA 6A
17 VIIA 7A
18 VIIIA 8A
H He Li Be B C N O F Ne Na Mg Al Si P S Cl Ar K Ca Sc Ti V Cr Mn Fe Co Ni Cu Zn Ga Ge As Se Br Kr Rb Sr Y Zr Nb Mo Tc Ru Rh Pd Ag Cd In Sn Sb Te I Xe Cs Ba Hf Ta W Re Os Ir Pt Au Hg Tl Pb Bi Po At Rn Fr Ra Rf Db Sg Bh Hs Mt Ds Rg Cn Nh Fl Mc Lv Ts Og
57-71
89-103
Lanthanide Series
La Ce Pr Nd Pm Sm Eu Gd Tb Dy Ho Er Tm Yb Lu
Actinide Series
Ac Th Pa U Np Pu Am Cm Bk Cf Es Fm Md No Lr
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B. Also know these ions:

Ag → _________

Zn → _________

III. Compounds That Contain Ions

A. Ions combine to form

____

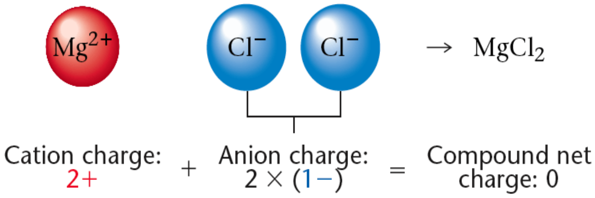
B. Usually composed of

___________ & ________________

IV. Formulas for ionic compounds

A. Write the _________ element symbol __________, followed by the ___________ element symbol.

B. The number of _____________ and ___________ must be correct for their ______________ to add up to _________



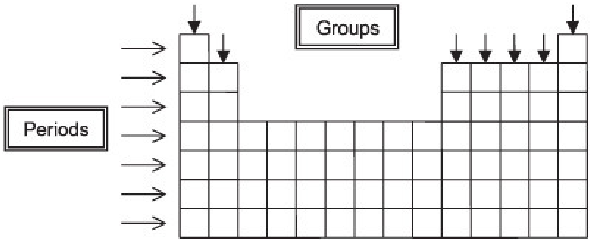
Ion & ionic compound practice problems:

Example #1:
Aluminum and Sulfur

Example #2:
Lithium and Nitrogen

Example #3:
Magnesium and Oxygen

I. Periodic Table Groups



A. Periods

1. Elements arranged in order of increasing ____________________

2. Arranged in ____

3. Labeled __________________________

B. Group/Family

1. Elements with similar ___________________________________

2. Arranged in a ____

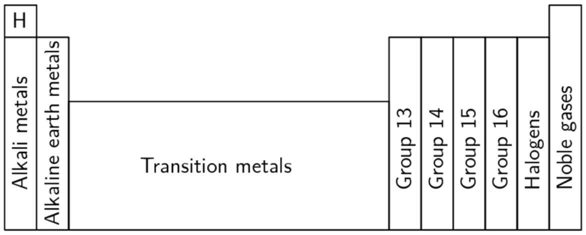
C. Locations of Major Groupings on the Periodic Table

1. Metals: appear in the _______________ and on the ___________ side.

2. ______________________: appear on the right side

3. ___________________________: show a blend of __________________ and nonmetal properties

   a. heavy line ____________________ separates metals and nonmetals



II. Properties of the Three Types of Elements

A. Metals

1. Efficient conduction of _______________ and ________________________

2. ___________________________ (can be hammered into thin sheets)

3. ______________________ (can stretch without breaking)

4. Lustrous (_______________) appearance

5. Most are _________________ at 25°C (room temperature) except _______ (a liquid)

B. Three Types of Metals

1. Alkali Metals (in the picture is sodium metal, which is soft and easy to cut. It reacts violently with many substances, including water)

   a. Group ____ on the periodic table

   b. Extremely __________________

      i. never found ____________ in ________________________

   c. Reactivity of these elements increases __________________ the group

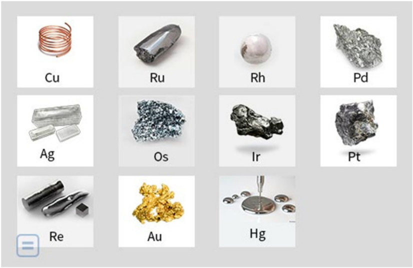
2. Alkaline Earth Metals (in the picture at right is calcium metal. Calcium metal also reacts with water, but far less violently)

   a. Group ___ on the periodic table

   b. Also ____

      i. never found _________________ in ____________________

3. ____



   a. All _______________ outside groups______ and _______

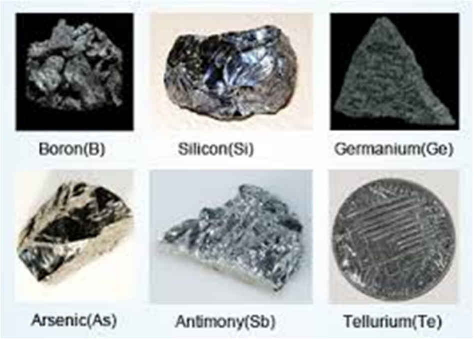
   b. Much more _________________ and less _________________________

   c. Some are found ________________ in nature

C. Metalloids/semimetals

1. Show a blend of _______________ and nonmetal __________________

2. Example: ______________________ is shiny like a metal but brittle and does not conduct ___________________ well



D. Nonmetals

1. Lack the properties of metals

2. Various states of matter at 25°C.

Examples:

Gas: ____________________

Liquid: __________

Brittle solid: _________________

E. Two of the non-metal groups are especially important

1. ____________________________ (example: yellow chlorine gas pictured at right)

a. Highly _____________________

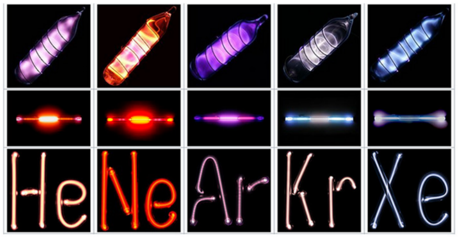
b. ____________ found ______________ in nature

2. ____________________________

a. ___________________________

b. Found in nature only as _________________________

c. ___________________ and ______________________



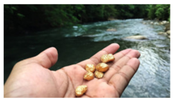
III. Natural States of the Elements

A. Most elements are very _______________

1. Not generally found in __________, uncombined form _______________

2. Exceptions are:

a. Noble metals – ____________________________



(this is why you can find gold nuggets in streams)

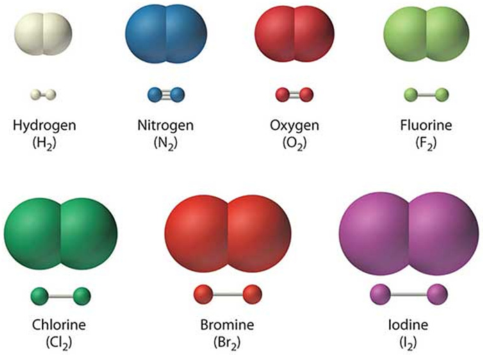
b. _______________________________ – Group 18

B. Diatomic Molecules

1. When pure, some elements exist only as molecules of ________ atoms

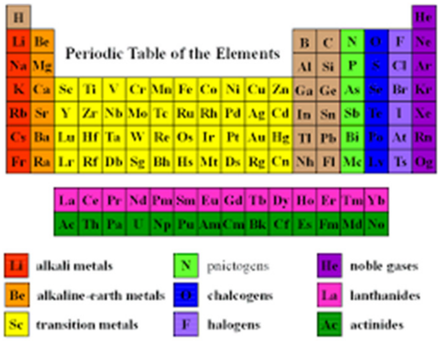
2. "Super seven": ____

(Just remember "HOFBrINCl")



Periodic Table Practice

Element	Symbol	Atomic #	Metal/Metalloid/ Nonmetal	Family
	Ar			Noble gases
Chlorine				
		56		
Cesium				



I. Periodic Table Trends

A. Trends are patterns

1. There are patterns for many properties of atoms

2. Example pattern: number of protons tends to ___________________

as you move ______________ and to the _____________

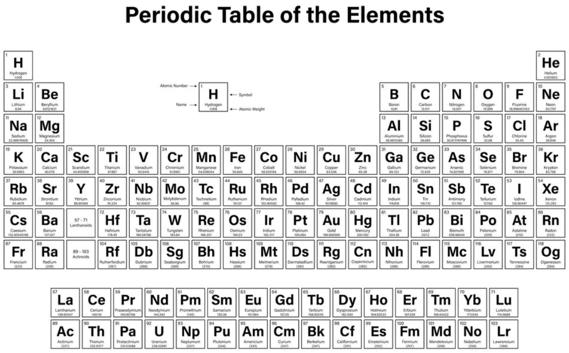
3. Example questions:

a. Which has more protons: copper or zinc?

i. _____________ has more protons because it is further ____________________

b. Which has more protons: copper or silver?

i. _____________ has more protons because it is further ____________________


Periodic Table of the Elements

II. Another example of a periodic table trend: atomic radius

A. No two atoms are the _______________________

1. Reminder: most of an atom's volume is its __________________________

B. Two competing factors determine size

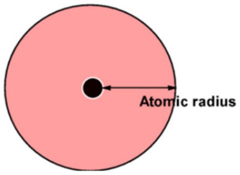
1. # of electrons: more electrons =

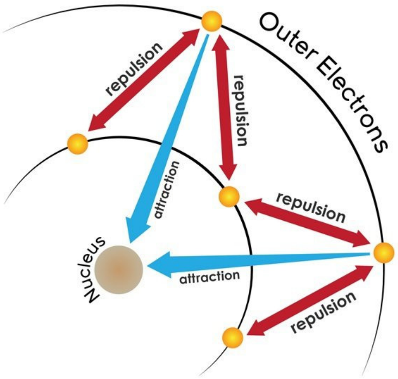
_____________ electron cloud

2. # of _______________: more ________________ =

more force pulling electrons ________________

(________________ electron cloud)





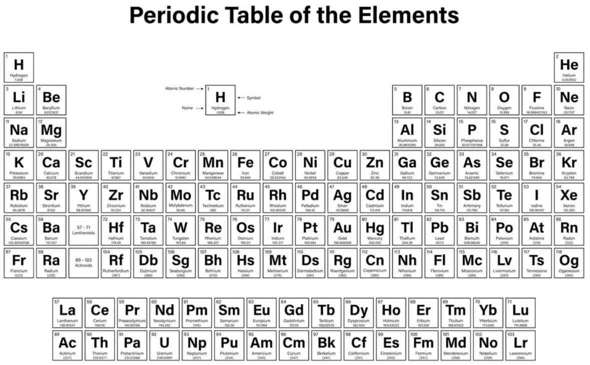
C. Net result: atomic size increases as you move

__________________ and toward the ________________

on the periodic table

(label in the box and draw arrows on the periodic table)

Trend for ___________________________________


Periodic Table of the Elements

D. Practice question: which element has a greater atomic

radius: Na or ______?

a. Answer: _____ because overall the trend says that it

gets bigger towards the _________ and _________, and

______ is overall closer to that corner of the periodic table

II. First Ionization energy

A. The amount of energy needed to remove the outermost ________________

1. Low ionization energy = ___________ to remove _____________

a. Easily becomes a __________________

b. Highly __________________

c. Usually a _____________

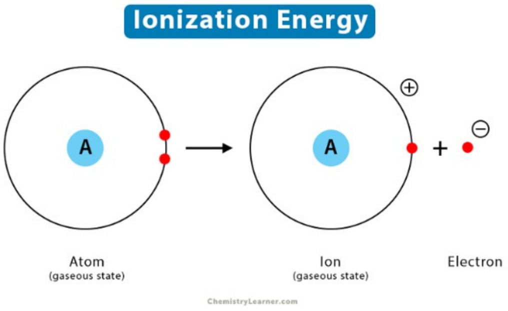
2. High ionization energy = _________________ to remove _________________

a. Many are less ___________

b. Many are _______________, which tend to gain ___________________

B. Atoms with a smaller atomic ________________ tend to have greater __________________

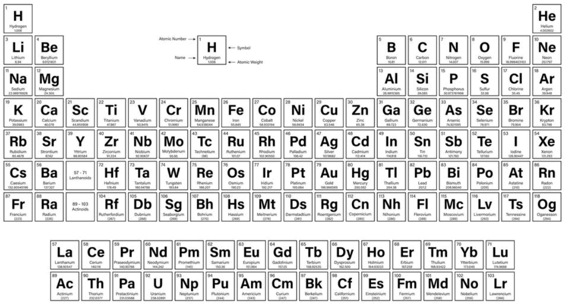
forces holding the _________________ to the ________________



C. Net result: first ionization energy increases as you move _________________ and toward the _______________ on the periodic table (label in the box and draw arrows on the table to the right)

This periodic table shows the trend for
____

Periodic Table of the Elements



H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	57 - 71 Lanthanoids	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	89 - 103 Actinoids	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og

La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

D. Practice question: which element has a greater first ionization energy: N or _____?

a. Answer: _______ because overall the trend says that it gets bigger towards the _________ and ________, and _______ is overall closer to that corner of the periodic table

III. Electronegativity

A. Tendency of an atom to attract __________________ toward itself

B. Atoms with high ionization energies tend to be highly electronegative

C. Net result: electronegativity increases as you move _________________ and toward the _______________ on the periodic table (label in the box and draw arrows on the table below)

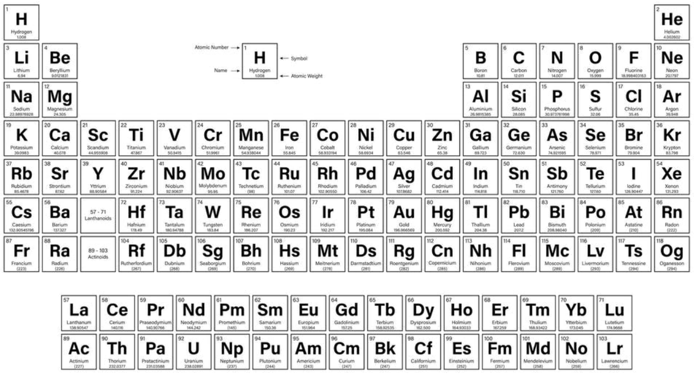
Fill this in

Electronegativity

Power to _______________________

Trend for ____________________________

Periodic Table of the Elements



H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	57 - 71 Lanthanoids	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	89 - 103 Actinoids	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og

La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

D. Practice question: which element has a greater electronegativity: I or ______?

a. Answer: _____ because overall the trend says that it increases towards the __________ and ____________, and _____ is overall closer to that corner of the periodic table

Name: ______________________ Per: ____

Note Questions: Atomic Structure and the Periodic Table

Notes #4: Atomic Structure & Isotopes

1. The three subatomic particles:
    a. Electrons: what is their charge? ________ Are they located inside or outside the nucleus? _____
    b. Neutrons: what is their charge? ________ Are they located inside or outside the nucleus? _____
    c. Protons: what is their charge? ________ Are they located inside or outside the nucleus? _____

2. Atomic number represents the number of ______________ and the mass number represents the number of ______________________

3. The cathode ray tube led to the discovery of what?

4. How did JJ Thompson determine that the particles in the cathode ray tube were negatively-charged?

5. Did the different elements tested in the cathode ray tube produce same or different particles?

6. Draw the plum-pudding model that resulted from the cathode ray experiment:

7. In the gold foil experiment, did most of the alpha particles pass through or bounce off the foil?

8. What were the three conclusions for this experiment?
    a.
    b.
    c.

9. Is the number of protons different or the same between two isotopes of an element? __________ What about the number of neutrons, same or different between two isotopes of the same element? __________

10. How many protons, neutrons, and electrons in neutral iron-58?
    Protons: _______ Neutrons: _______ Electrons: _______

11. What is the name (see example in previous question) of a neutral atom with 20 protons and 21 neutrons?

Notes #5: Ions & Ionic Compounds

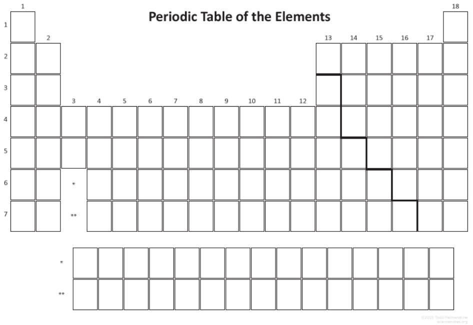
12. Atoms can form ions by gaining or losing what?

13. Positive ions are called what? ___________ Name two examples of positive ions: _______ & _______

14. Negative ions are called what? ___________ Name two examples of negative ions: _______ & _______

15. Do negative ions or positive ions go first in a chemical formula?

16. Label the charge of the ions at the top of each column:



17. Draw a diagram showing how to "cross" the charges and find the formula for the following compounds:
a. Lithium and chlorine: b. Oxygen and magnesium

Notes #6: The Periodic Table

18. What is a "period" on the periodic table?

19. What is a "group" (AKA "family") on the periodic table?

20. List the three major groupings of elements:
a. ________________ b. ________________ c. ________________

21. List five properties of metals:
a. __________________ c. __________________ e. __________________
b. __________________ d. __________________

22. What is group 1 called?

23. Describe the chemical properties of group 1

24. What is group 2 called?

25. Describe the chemical properties of group 2

26. What is group 17 called?

27. Describe the chemical properties of group 17

28. What is group 18 called?

29. Describe the chemical properties of group 18

30. What are the metals in the middle of the periodic table usually called?

31. Are most elements found pure in nature?

32. List the seven diatomic elements:
a. __________________ c. __________________ e. __________________ g. __________________
b. __________________ d. __________________ f. __________________

Name: ____________________________ Per: ____

Worksheet: Practice with Ions

1. List the charges of the subatomic particles:

Proton: ___________________ Neutron: ___________________ Electron: ___________________

2. Define what an ion is: ____

For 3-5, use fill the blanks with the names of subatomic particles

3. If you change the number of __________________ in an atom, it becomes a different element

4. If you change the number of __________________ in an atom, it becomes an ion of the same element

5. If you change the number of __________________ in an atom, it does not become an ion and is a different isotope of the same element

6. Adding an electron makes an atom more ____________________, whereas subtracting an electron makes an atom more _________________ (hint: put the term "positive" or "negative" in each blank)

7. Write the term that means "positive ion": ______________ and the term that means "negative ion": ______________

8. Use the chart to express how to create the desired charge on an atom

Change in number of electrons	Charge of the resulting ion
Subtract 3 electrons	Charge of +3
Add 2 electrons	Charge of -2
Subtract 1 electron	Charge of +1
Add 5 electrons	
Subtract 1 electrons	
Add 2 electrons	
Subtract 3 electrons	
	Charge of +7
	Charge of -2
	Charge of +1

9. Use the periodic table to predict the charge for each element when it becomes an ion, and to explain how many electron(s) it gained or lost in order to have that charge

Element	Charge of ion	Formula of ion	How it got that charge
Aluminum	+3	Al^{3+}	Lost three electrons
Sodium	+1	Na^{+}	Lost one electron
Fluorine	-1	F^{-}	Gained one electron
Oxygen	-2	O^{2-}	Gained two electrons
Nitrogen			
Magnesium			
Potassium			
Sulfur			
Bromine			
Aluminum (yes, this repeat is intentional)			

Be sure to complete all questions on both sides of this paper!

10. You will need to memorize the charges for these ions:
Silver: ____________________________
Zinc: ____________________________
Aluminum: ____________________________

11. What kind of ion (cation or anion) always has “-ide” in its name? ___________________________________

12. Name the ion (see notes or use google to check your answers if unsure!)

Element	Formula (as an ion)	Name
Sodium	Na^+	Sodium ion
Nitrogen	N^{3-}	Nitride ion
Hydrogen*	H^+	
Hydrogen	H^-	
Lithium		
Oxygen		
Chlorine		
Magnesium		
Aluminum		
Silver		
Zinc		
Sulfur		
Fluorine		

*Yes, hydrogen can be both negative and positive in charge

13. How many protons, neutrons and electrons for each?

Name	Number of Protons	Number of Neutrons	Number of electrons
Fe^{+3}	26	56 – 26 = 30*	26 – 3 = 23
Sulfur-31 with a charge of -2	16	31 – 16 = 15	16 + 2 = 18
Calcium-39 with a charge of +2			
Zn^{+2}			
Cl^-			
Phosphorus-32 with a charge of -3			

*Reminder – if not given a specific isotope of an element, then use the rounded mass number from the periodic table

Name: ____________________________ Per: ____

Worksheet: Practice with Ionic Compounds

1. Write the term that means “positive ion”: ___________________ and the term that means “negative ion”: ___________________

2. In an ionic compound, the charges of the ions involved must add up to what number? _______

For 3, use the terms “positive” and “negative”
3. In an ionic compound, the ion that goes first is always ______________, and the one that goes last is always ______________

4. Predicting the formulas of ionic compounds:

Elements	Write them in the correct order with their charges	Switch the charges	Correct final formula
Sulfur and Aluminum	$Al^{3+}S^{2-}$	Al_2S_3	Al_2S_3
Potassium and Nitrogen	K^+N^{3-}	K_3N_1	K_3N
Sodium and Fluorine	Na^+F^-	Na_1F_1	NaF
Beryllium and Nitrogen			
Iodine and Lithium			
Barium and Sulfur			
Phosphorus and Aluminum			
Strontium and Fluorine			
Bromine and Aluminum			
Phosphorus and Sodium			

7. The formula Ca_1Cl_2 is wrong.
   a. Explain why the formula is wrong:

   b. Write the correct formula: ________________

8. The formula ClK is wrong.
   a. Explain why the formula is wrong:

   b. Write the correct formula: ________________

9. The formula Mg_2O_2 is wrong.
   a. Explain why the formula is wrong:

   b. Write the correct formula: ________________

10. The formula Ca_2N_3 is wrong.
    a. Explain why the formula is wrong:

    b. Write the correct formula: ________________

Name: ________________________________ Per: ____

Worksheet: Practice with Isotopes and Atomic Structure

1. The nucleus of the atom is composed of ________________ and ________________

2. The _____________________ of the atom are found outside the nucleus

3. Isotopes of a given element have the same number of _______________________ and differing numbers of _____________________ (insert the terms "electrons", "protons", and "neutrons" into these blanks. Yes, that means one of the blanks contains two terms)

4. The mass number of an atom represents the number of ________________________

5. The atomic number of an atom represents the number of ________________________

6. For each of the following elements, use a periodic to write the atomic number and average atomic mass number. See first row of the chart for a completed example.

Element Symbol	Element Name	Atomic Number	Average Atomic Mass
Si	Silicon	14	28.086
Sr			
	Cobalt		
	Copper		
	Carbon		
B			
Be			
	Lead		
	Mercury		
		26	
			4.003

7. Give the atomic mass, number of protons and number of neutrons for an AVERAGE atom for each of the following elements. Show how you calculated the number of neutrons. See first row of the chart for a completed example.

Element	Atomic Mass (Raw)	Atomic Mass (Rounded to whole number)	Number of Protons	Number of Neutrons
Aluminum	26.982	27	13	27 – 13 = 14
Iron				
Hydrogen				
Tin				
Lead				

Be sure to complete all questions on both sides of this paper!

8. Complete the chart, showing how you calculated the number of neutrons for each. See first row of the chart for a completed example.

Name	Number of Protons	Number of Neutrons	Atomic Number	Atomic Mass (for this specific isotope)	Symbol
Iron-57	26	57 – 26 = 31	26	57	${}^{57}_{26}\text{Fe}$
Boron-10					
Iodine-131					
Hydrogen-3					
					${}^{67}_{29}\text{Cu}$
					${}^{40}_{19}\text{K}$
					${}^{294}_{118}\text{Og}$

9. Write the atomic symbol using the format ${}^{Y}_{Z}X$ for each of the isotopes described below.

Example: An atom with three protons and four neutrons:

Name: Lithium-7 Symbol: ${}^{7}_{3}\text{Li}$

a. A nitrogen atom with 7 protons and 8 neutrons:

Name: ___________________ Symbol: ___________________

b. An element with 26 protons and 31 neutrons

Name: ___________________ Symbol: ___________________

c. The isotope of iodine with a mass number of 131

Name: ___________________ Symbol: ___________________

d. The isotope with 25 protons and a mass number of 54

Name: ___________________ Symbol: ___________________

e. The isotope of chlorine in which mass number is 37

Name: ___________________ Symbol: ___________________

f. Cobalt with 33 neutrons

Name: ___________________ Symbol: ___________________

g. An element with 21 neutrons and a mass number of 41

Name: ___________________ Symbol: ___________________

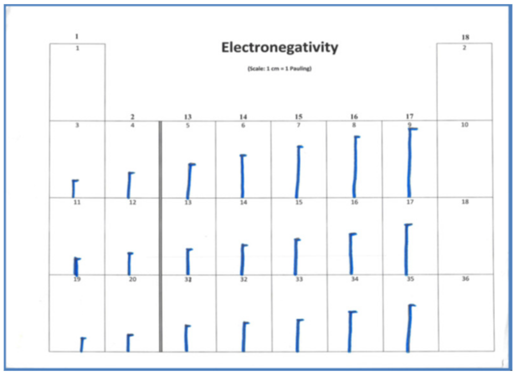
h. An element with 1 proton and zero neutrons

Name: ___________________ Symbol: ___________________

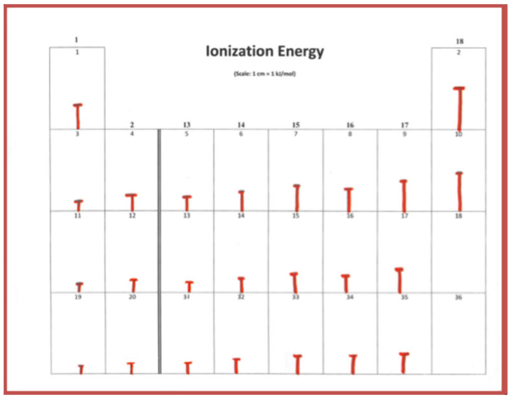
Name: ________________________ Per: ____

Introduction to Trends of the Periodic Table

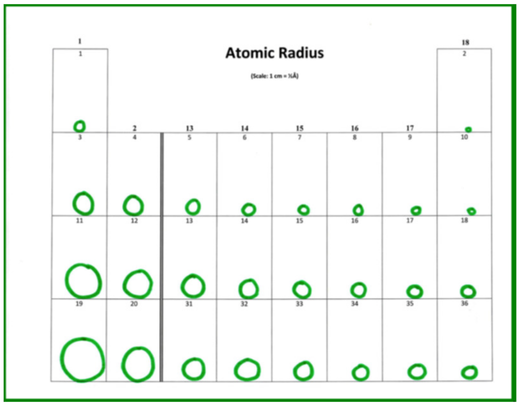
1. Name of trend: ______________________________________
    a. Is the trend increasing as you go up or down a group? _______
    b. Is the trend increasing as you go left or right across a period? _______
    c. This trend in _________________ is increasing as you go ______ and to the __________
    d. Draw arrows to show this trend on the right and bottom sides of the diagram.



2. Name of trend: ______________________________________
    a. Is the trend increasing as you go up or down a group? _______
    b. Is the trend increasing as you go left or right across a period? _______
    c. This trend in _________________ is increasing as you go ______ and to the ______
    d. Draw arrows to show this trend on the right and bottom sides of the diagram.



3. Name of trend: ______________________________________
    a. Is the trend increasing as you go up or down a group? _______
    b. Is the trend increasing as you go left or right across a period? _______
    c. This trend in _________________ is increasing as you go ______ and to the __________
    d. Draw arrows to show this trend on the right and bottom sides of the diagram.

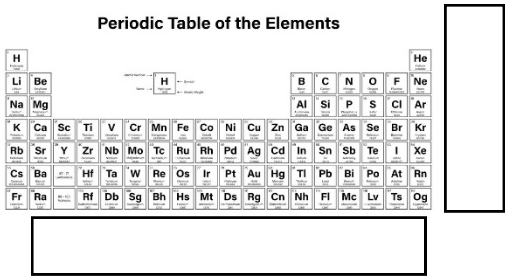


Name: ________________________ Per: _____

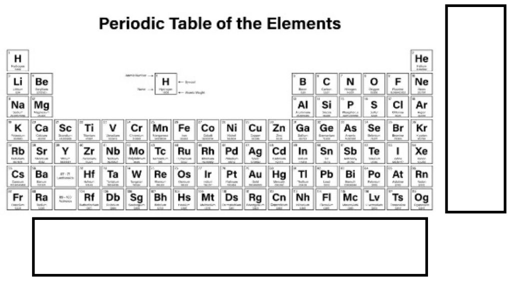
Worksheet: Periodic Table Trends Practice

Instructions: For each trend, you will draw the trend in the boxes to the right/bottom of the periodic table. Draw an arrow in each box pointing left/right and up/down in order to show the direction in which the trend increases. Example: if it increases as you go up, then insert an up arrow, and if it increases as you go downwards, then insert a down arrow into the box.

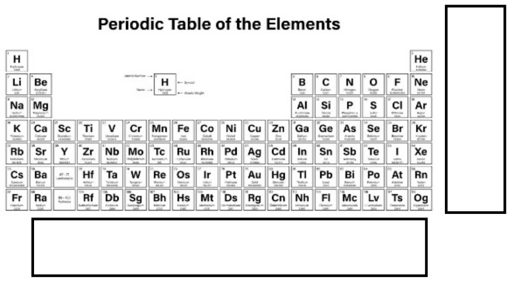
1. Atomic radius: place the appropriate arrows in the boxes for the directions in which this INCREASES


Periodic Table of the Elements
H He
Li Be H B C N O F Ne
Na Mg Al Si P S Cl Ar
K Ca Sc Ti V Cr Mn Fe Co Ni Cu Zn Ga Ge As Se Br Kr
Rb Sr Y Zr Nb Mo Tc Ru Rh Pd Ag Cd In Sn Sb Te I Xe
Cs Ba Hf Ta W Re Os Ir Pt Au Hg Tl Pb Bi Po At Rn
Fr Ra Rf Db Sg Bh Hs Mt Ds Rg Cn Nh Fl Mc Lv Ts Og

2. Ionization energy: place the appropriate arrows in the boxes for the directions in which this INCREASES


Periodic Table of the Elements
H He
Li Be H B C N O F Ne
Na Mg Al Si P S Cl Ar
K Ca Sc Ti V Cr Mn Fe Co Ni Cu Zn Ga Ge As Se Br Kr
Rb Sr Y Zr Nb Mo Tc Ru Rh Pd Ag Cd In Sn Sb Te I Xe
Cs Ba Hf Ta W Re Os Ir Pt Au Hg Tl Pb Bi Po At Rn
Fr Ra Rf Db Sg Bh Hs Mt Ds Rg Cn Nh Fl Mc Lv Ts Og

3. Electronegativity: place the appropriate arrows in the boxes for the directions in which this INCREASES


Periodic Table of the Elements
H He
Li Be H B C N O F Ne
Na Mg Al Si P S Cl Ar
K Ca Sc Ti V Cr Mn Fe Co Ni Cu Zn Ga Ge As Se Br Kr
Rb Sr Y Zr Nb Mo Tc Ru Rh Pd Ag Cd In Sn Sb Te I Xe
Cs Ba Hf Ta W Re Os Ir Pt Au Hg Tl Pb Bi Po At Rn
Fr Ra Rf Db Sg Bh Hs Mt Ds Rg Cn Nh Fl Mc Lv Ts Og

4. Elements practice: find a periodic table and use it (and your knowledge of the trends) to identify which elements has the greater trend

Cobalt and Zirconium (Zr, #40) a) Which is more electronegative? _____________ b) Which has a greater ionization energy? _______ c) Which has a greater atomic radius? __________	Carbon and Fluorine a) Which has a LESSER atomic radius? ___________ b) Which has a GREATER ionization energy? ______ c) Which is LESS electronegative? _______________
Lithium and Beryllium a) Which has a greater atomic radius? __________ b) Which is more electronegative? _____________ c) Which has a greater ionization energy? _______	Tellurium (Te, #52) and Chlorine a) Which is less electronegative? ________________ b) Which has a lesser ionization energy? __________ c) Which has a lesser atomic radius? _____________

5. Trends practice: organizing multiple elements
   a. Consider the elements Rb (#37), V(#23), Al, and F.
      i. Which has the GREATEST atomic radius? _____ Which has the SMALLEST atomic radius? _____
      ii. Which has the MOST ionization energy? _____ Which has the LEAST ionization energy? _____
      iii. Which is the MOST electronegative? _____ Which is the LEAST electronegative? _____

Topic List: Unit 3 (Atomic Structure and the Periodic Table)

The following is a list of topics, concepts, and/or specific tasks that you are expected to have mastered. DO NOT RELY SOLELY ON THIS DOCUMENT!! This is NOT a study guide, nor is it in any way a substitute for the study guide that you should be making for yourself. There may be topics included on the exam that are not on this list. Use this list as a general starting place in your studying for the exam.

Here is a list of topics that you should understand and be able to define, explain, etc.

Notes #4: Atomic Structure & Isotopes

1. Structure of the atom
2. Atomic mass and atomic number
3. Dalton's atomic theory
4. Cathode ray experiment – method, results, findings, etc
5. Gold foil experiment – method, results, findings, etc
6. Plum pudding model of the atom: what it showed, what evidence it was based on, and how it was proved wrong
7. Isotopes
8. Isotope names (e.g. uranium-235) and their symbols, e.g. $^{235}_{92}\mathrm{U}$
9. Calculating how many neutrons, protons, etc in any isotope, charged or neutral

Notes #5: Ions & Ionic Compounds

10. What an ion is, and how it is formed from a neutral atom
11. How to name ions
12. Determine the charge of ions
13. Determine the formula of any ionic compound

Notes #6: The Periodic Table

14. Periods vs groups vs families
15. Names, properties of the different chemical groups
16. Metals vs metalloids vs non-metals
17. Properties of Metals
18. Whether elements in a group are found pure in nature or not
19. Diatomic elements

Notes #7: Trends of the Periodic Table

20. You may be asked to memorize the trends (ask your teacher). If so, then know them.
21. The trends dealt with on the exam are atomic radius, first ionization energy and electronegativity. Know what each of those terms means.
22. Be able to use the trends to make predictions about elements. For example, is given a choice between two elements (sodium and chlorine for example), you should be able to tell which is more electronegative and which is less electronegative, which has greater ionization energy or lesser ionization energy, and which has a greater or lesser atomic radius.

Unit 4: Modern Atomic Theory

I. Review: structure of the atom

A. Tiny nucleus composed of _____________ and _________________

1. Most of the mass is in the __________________

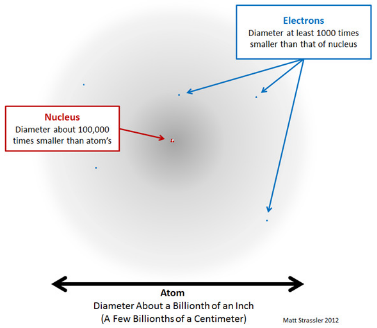
B. Electrons outside _________________

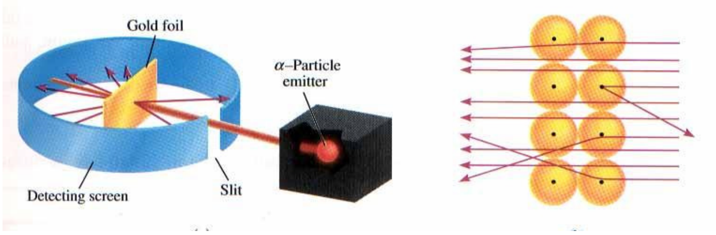
1. Electron cloud makes up most of the atom’s ____________________

II. Rutherford’s Gold Foil Experiment

A. Rutherford demonstrated that the atom has a ________________,

dense _____________ with his gold foil experiment


Electrons
Diameter at least 1000 times smaller than that of nucleus
Nucleus
Diameter about 100,000 times smaller than atom’s
Atom
Diameter About a Billionth of an Inch
(A Few Billionths of a Centimeter)
Matt Strassler 2012



B. Rutherford thought that the electrons were ________________________ around the atom, and all the __________________

Draw Rutherford’s early model of the atom here:

Problem with Rutherford’s model:

Question: If all electrons are the same, then why do different elements give off different light when energized with electricity?

They should look the ____________

Answer: Rutherford’s assumption that the electrons are ___________________________ through the atom was __________

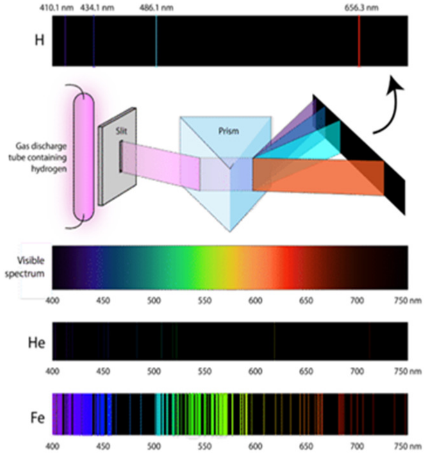
III. Line emission spectrum

A. This is the natural phenomenon that proved Rutherford’s early model wrong

1. “Spectrum” = the _________ you get when white light goes through a prism

B. Each element only produces some of the colors, so instead of a full rainbow, you see only ___________ of color

C. Each ______________ has its own unique “fingerprint” of color lines

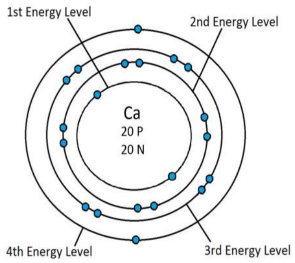


IV. Bohr Model

A. Niels Bohr, working with Rutherford, conducted experiments that showed that the electrons in an atom exist in different ____________________

B. The Bohr model shows electrons orbiting the nucleus like ____________________ around the ____________

Draw the Bohr model of the atom here:

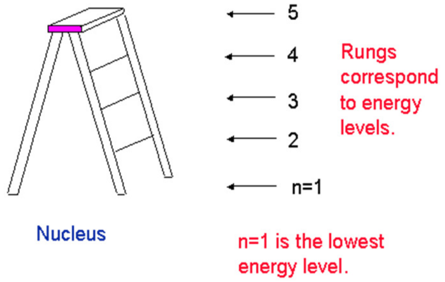


IV. Quantum energy levels

A. Electrons are divided into different ________________________

B. The energy levels are "quantum" levels with ________________ in between

Analogy: quantum energy levels are like a _____________: you can go up or down, but can't be in between!

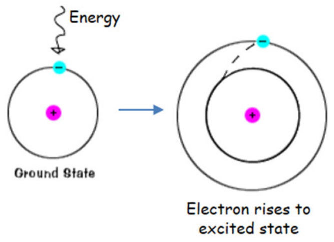


Rungs correspond to energy levels.

n=1 is the lowest energy level.

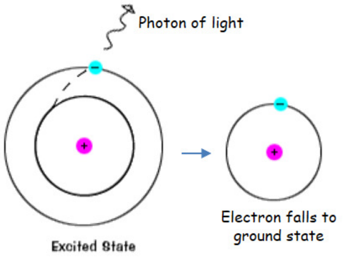
C. Four steps of how electrons in atoms give off light

1. Electrons absorb ____________________



2. Energy absorbed causes ______________ to rise to a ______________________________

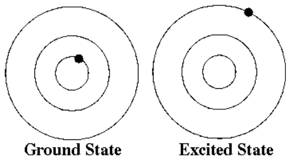
3. Electron needs to _______ back down to _______________ energy level



4. In order to fall back down to original energy level, electron must ______________ excess ______________ as ________________

D. Vocabulary:

1. "Excited state" = _____________ energy level
2. "Ground state" = _____________ energy level


Ground State Excited State

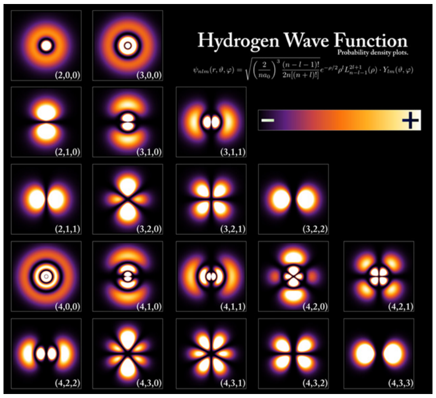
E. Because different elements have different _________________ of electrons, each produces its own ______________ blend of ________________

V. Wave Mechanical Model

A. Bohr was partially correct; his predictions only worked for _________________

B. Bohr's ideas about __________________________ were correct

C. Schrödinger and DeBroglie: electrons have properties of both a _____________ and a ________________

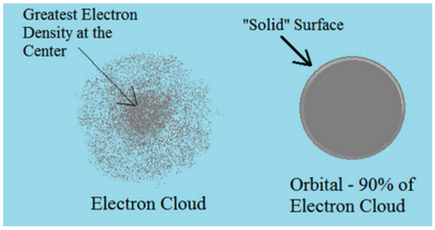


1. Electrons exist in _________________

   a. _________________ are NOT orbits

   b. The electron does not travel like a particle in simple directions

   c. Orbitals are the ________________ where an __________ is likely to be _________ of the time

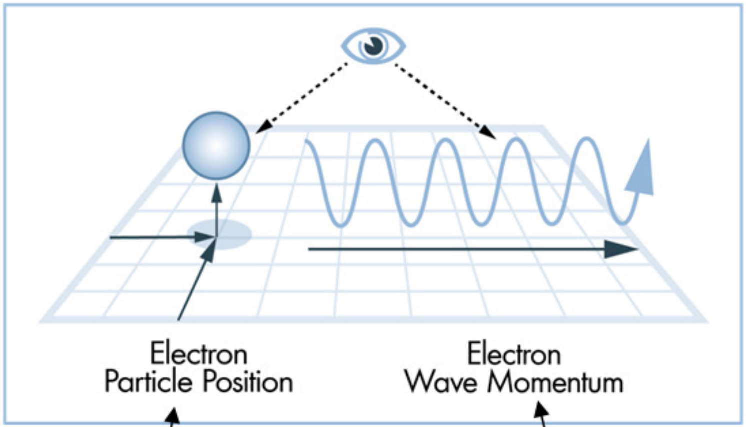

Greatest Electron Density at the Center
"Solid" Surface
Electron Cloud
Orbital - 90% of Electron Cloud

Heisenberg's Uncertainty Principle...

You can find out where the electron is going, but not where it is!

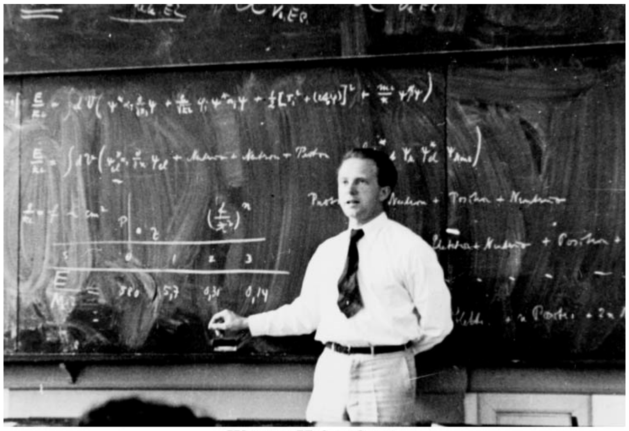
OR…

You can find out where the electron is, but not where it is going.



This is the basis of the points plotted on a 3-D graph to represent an orbital

This is the basis of the idea that you never know where the electron is when it's moving inside the orbital. It is spread out, simultaneously everywhere and nowhere at the same time!



Werner Heisenberg

I. Electron Orbitals

A. The _____________ where you are most ________________ to find an ____________

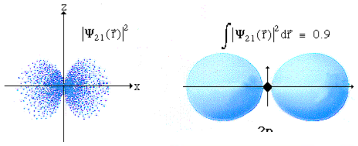
1. Contains _____ of the possible ___________where an ________________ might be found

B. Electron _______________________________

1. Symbolized by "____" (the Principal Quantum Number)

2. Related to the probable ___________________ of the __________________ from the ___________________

a. Higher energy level = __________________ distance from the _____________________



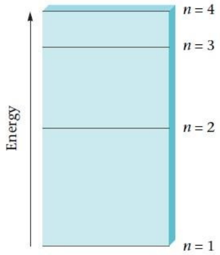


Figure 11.13 The first four principal energy levels in the hydrogen atom. Each level is assigned an integer, *n*.

II. Electron energy level characteristics

A. Higher-energy ____________________ can carry more ____________________

B. Number of __________________ in each energy level is ______________, where _____ is the principal quantum _____________

n = 1 has ____ electrons, n = 2 has ____ electrons, n = 3 has ____ electrons, etc

C. Each _______________ has ___________________

ex. in hydrogen atoms:

Level 1 has ___ sublevel

Level 2 has ___ sublevels

Level 3 has ___ sublevels

Level 4 has ___ sublevels

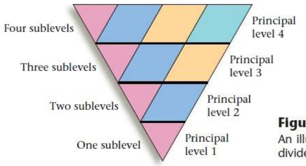


Figure 11.14 An illustration of how principal levels can be divided into sublevels

III. Sublevel Types

A. Each energy level contains ___________________________

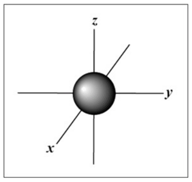
1. _____________________

a. Found at ______ energy levels

b. _________________________

c. Holds up to ________ electrons

d. Contains _______ orbital



e. Each orbital contains up to ___ electrons (____ total in the sublevel)

2. ______________________

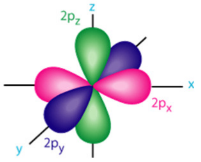
a. Found at energy levels ____ and up

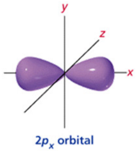
b. _______ different kinds, all _____________ shaped

c. Holds up to ________ electrons

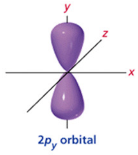
d. Contains _______ orbitals

e. Each orbital contains up to ___ electrons (____ total in the sublevel)

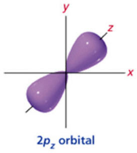




$2p_x$ orbital



$2p_y$ orbital



$2p_z$ orbital

3. _____________________

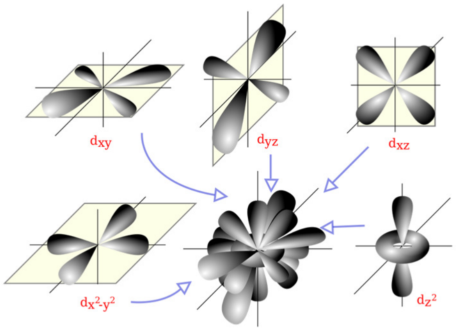
a. Found at energy levels ____ and up

b. _________ different kinds, shaped like dumbells and donuts

c. Holds up to ________ electrons

d. Contains _______ orbitals

e. Each orbital contains up to ___ electrons (____ total in the sublevel)



4. ____________________

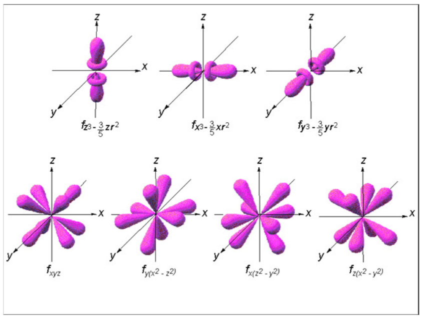
a. Found at energy levels ____ and up

b. _________ different kinds, shaped like dumbells and donuts

c. Holds up to ________ electrons

d. Contains _______ orbitals

e. Each orbital contains up to ___ electrons (____ total in the sublevel)


z
x
y
$f_{z^3-\frac{3}{5}zr^2}$
$f_{x^3-\frac{3}{5}xr^2}$
$f_{y^3-\frac{3}{5}yr^2}$
f_{xyz}
$f_{y(x^2-z^2)}$
$f_{x(z^2-y^2)}$
$f_{z(x^2-y^2)}$

Energy Levels, Orbitals, Electrons

Energy Level (n)	Orbital type in the energy level (AKA sublevels)	Number of orbitals	Number of Electrons	Number of electrons per Energy level ($2n^2$)
1	s	1	2	2
2	s p	1 3	2 6	8
3	s p d	1 3 5	2 6 10	18
4	s p d f	1 3 5 7	2 6 10 14	32

IV. Filling the Orbitals

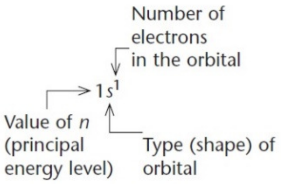
A. Electrons will always fill the _____________ energy orbitals _____________

B. _____________________ atoms don’t have enough ______________________ to fill the outer ________________________________

ex. H has electrons in level ____, but Mg has electrons in level _____

I. Electron Configurations

A. Describes how the ________________ are spread out among the different __________________ of an ___________


Number of electrons in the orbital
$1s^1$
Value of n (principal energy level)
Type (shape) of orbital

B. Example: ______________

1. Electron configuration: __________

- this means _______ electron in the ____ orbital of the _____ energy level

C. Example #2: _____________

1. Electron configuration: _________

- ______ electrons, both in the ______ orbital

II. Electrons fill ______________ in a certain order

A. Order of filling:

First: electrons fill level _____ (has ____ orbital only)

Next: electrons fill level _____ (has ____ and _____ orbitals)

Next: electrons then fill levels _________________________

B. Examples:

Hydrogen: ___ electron __________________

Helium: ___ electrons __________________

Lithium: ___ electrons __________________

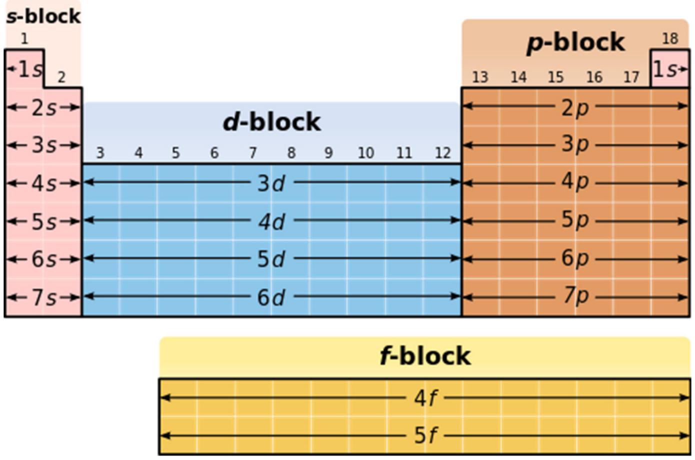
Beryllium: ___ electrons __________________

Boron: ___ electrons __________________

C. Electron Configuration Chart

1s
2s 2p
3s 3p 3d
4s 4p 4d 4f
5s 5p 5d 5f
6s 6p 6d
7s 7p
8s

The periodic table can also be used to predict electron configuration



Example electron configurations:

Analyze these for the purpose of understanding why they are what they are…

- Lithium: $1s^2 2s^1$
- Oxygen: $1s^2 2s^2 2p^4$
- Potassium: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
- Iron: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
- Krypton (#36): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
- Indium (#49): $11s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^1$
- Radon (#86): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6$

Practice

- Nitrogen: ____
- Calcium: ____
- Titanium: ____
- Arsenic: ____
- Argon: ____
- Lead: ____

C. Electron Configurations for Ions:

1. Add/subtract the appropriate number of electrons to/from the electron configuration

2. Examples:

Hydrogen: 1 electron	$1s^1$	Beryllium: 4 electrons	$1s^2 2s^2$
Hydrogen ion: ____ electrons	________________	Be^{2+} ion: ____ electrons	________________
Lithium: 3 electrons	$1s^2\ 2s^1$	Boron: 5 electrons	$1s^2 2s^2 2p^1$
Li^+ ion: ____ electrons	________________	B^{3+}: ____ electrons	________________

3. More examples:

Nitrogen: 7 electrons $1s^22s^22p^3$

N^{3-} ion: ____ electrons _________________

Oxygen: 8 electrons $1s^22s^22p^4$

O^{2-} ion: ____ electrons _________________

Fluorine: 9 electrons $1s^22s^22p^5$

F^- ion: ____ electrons _________________

Neon: 10 electrons $1s^22s^22p^6$

Neon ion: ______________________________________

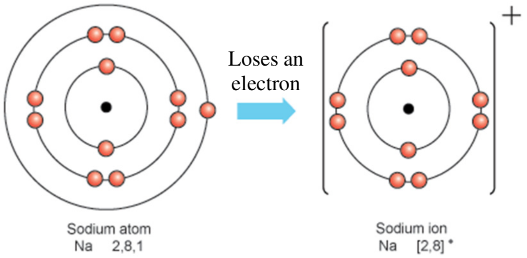
4. Atoms form ions because their electrons' outer _______________ level is _________________________

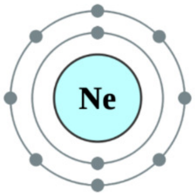
a. They will gain or lose the smallest possible number of _______________ in order to have a _________ outer energy level

b. Example #1: _________________ forms a +1 ion because it needs to ________ one electron so that it can have a _________ outer energy level

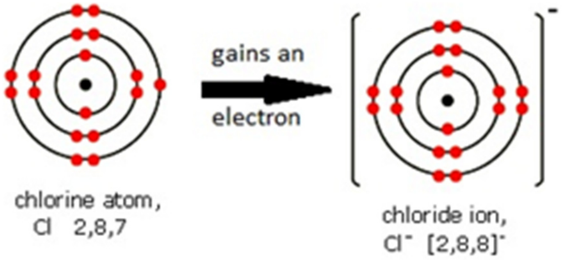
c. Example #2: _________________ forms a -2 ion because it needs to ________ two electrons so that it can have a _________ outer energy level

c. Example #3: neon does not form an ion because its ____________________ is already ________, so it does not need to gain or lose any __________





Noble gases such as neon neither gain nor lose electrons. Why? Answer in the space below:



4. Transition metal ions:

Remove electrons from the highest principal quantum number first!

Manganese: $1s^22s^22p^63s^23p^64s^23d^5$

Mn^{3+} ion: subtract 3 e^- ______________________

Cobalt: $1s^22s^22p^63s^23p^64s^23d^7$

Co^+ ion: subtract 1 e^- ______________________

III. A shorter way to write electron configurations

A. Electron configurations can get very long…

ex. ______________________: $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^1$

B. Shortcut: noble gas configurations

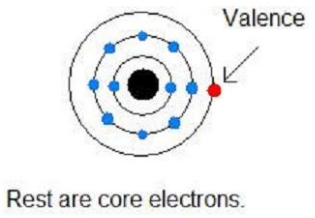
1. Just write the previous ___________________ and then write the extra _________________ which the element has

a. example: __________________ has _____ electrons

- The _____________ before _______________ is ________________ (54 electrons)
- _________________ is: $1s^22s2^2p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^6$
- _________________ is: $1s^22s2^2p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^1$
- Noble gas configuration for __________________ is: _______________

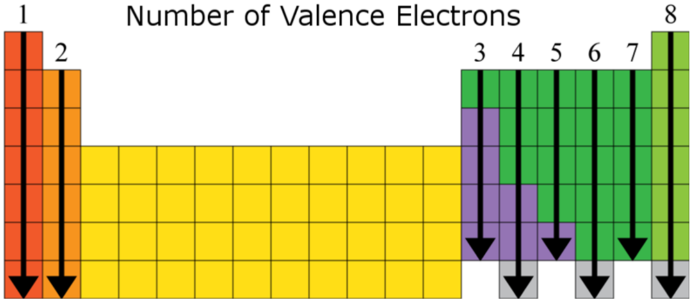
b. example #2: _______________ has _____ electrons

- Noble gas prior to _____________ is ________________ (_____ electrons)
- _____________ is: $1s^2 2s 2^2 p^6 3s^2 3p^6$
- _____________ is: $1s^2 2s 2^2 p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
- Noble gas configuration for __________________ is: _______________



C. Valence electrons

1. The ________________________ in an atom's electron shell
2. Participate in ______________________________________
3. Have the highest principal ___________________ number
   a. Example: [Xe] $6s^1$
      - this has ____ valence electron ________________________________
   b. Example #2: [Ne] $3s^2$ $3p^3$
      - this has ______ valence electrons _______________________________
   c. Example #3: [Ar] $4s^2$ $3d^5$
      - this has ______ valence electrons _______________________________



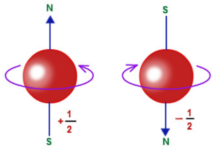
4. Number of valence electrons is the same in each _______________ (familiarize yourself with this pattern on the periodic table)

D. __________ electrons

1. All other _______________ that are not ________________ electrons
2. Do NOT participate in ____________________________

IV. Orbital Diagrams

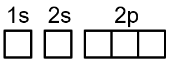
A. Rules for drawing orbital diagrams



1. Electrons ___________ and you need to show which way they are spinning
2. Pauli Exclusion Principle: "No two electrons can have the same four quantum numbers" (therefore, since an atomic orbital can hold up to _________ electrons, those electrons must have opposite ____________)
3. Hund's Rule: Every orbital in a sublevel is ________ occupied before any orbital is _____________ occupied
4. Aufbau Principle: electrons fill the _____________________ available energy levels before filling ____________________________________

B. Orbitals vs Sublevels

1. s, p, d, and f are called ___________________
2. Each sublevel contains ______________
3. Orbitals holds no more than _____ electrons each no matter where they are
   a. represented by boxes
      i. s sublevel has ___ orbital (max 2 electrons)
      ii. p sublevel has ___ orbitals (max 6 electrons)
      iii. d sublevel has ___ orbitals (max 10 electrons)
      iv. f sublevel has ___ orbitals (max 14 electrons)
4. Show orbitals as ___________ in your diagrams



S sublevels hold only 2 electrons because they contain one orbital each

P sublevels contain 3 orbitals, which is why they can fit up to 6 electrons

1. Hydrogen

$1s^1$

1s ☐

2. Helium

$1s^2$

1s ☐

3. Lithium – three electrons

$1s^2 2s^1$

1s ☐ 2s ☐

4. Beryllium – four electrons

$1s^2 2s^2$

1s ☐ 2s ☐

On this note handout, label one of the diagrams to show the difference between an orbital and a sublevel. Ask your instructor for clarification about how to do this!

5. Boron – five electrons

$1s^2 2s^2 2p^1$

1s ☐ 2s ☐ 2p ☐☐☐

6. Carbon – six electrons

$1s^2 2s^2 2p^2$

1s ☐ 2s ☐ 2p ☐☐☐

7. Nitrogen – seven electrons

$1s^2 2s^2 2p^3$

1s ☐ 2s ☐ 2p ☐☐☐

8. Oxygen – eight electrons

$1s^2 2s^2 2p^4$

1s ☐ 2s ☐ 2p ☐☐☐

9. Fluorine – nine electrons

$1s^2 2s^2 2p^5$

Draw your own diagram for fluorine

10. Neon – ten electrons

$1s^2 2s^2 2p^6$

Draw your own diagram for neon

11. Sodium – eleven electrons

$1s^2 2s^2 2p^6 3s^1$

1s ☐ 2s ☐ 2p ☐☐☐ 3s ☐

12. Titanium – 22 electrons

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$

V. Orbital diagrams of ions

1. Oxygen ion has -2 charge: ____ + ____ = ____ electrons

$1s^2 2s^2 2p^6$

1s ☐ 2s ☐ 2p ☐☐☐

2. Titanium 2+ ion: (____ - ____ = ____ electrons)

Neutral titanium: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$

always remove electrons from the _______________ principal quantum number

titanium ion configuration: ____________________________

Draw the orbital diagram for titanium 2+ ion:

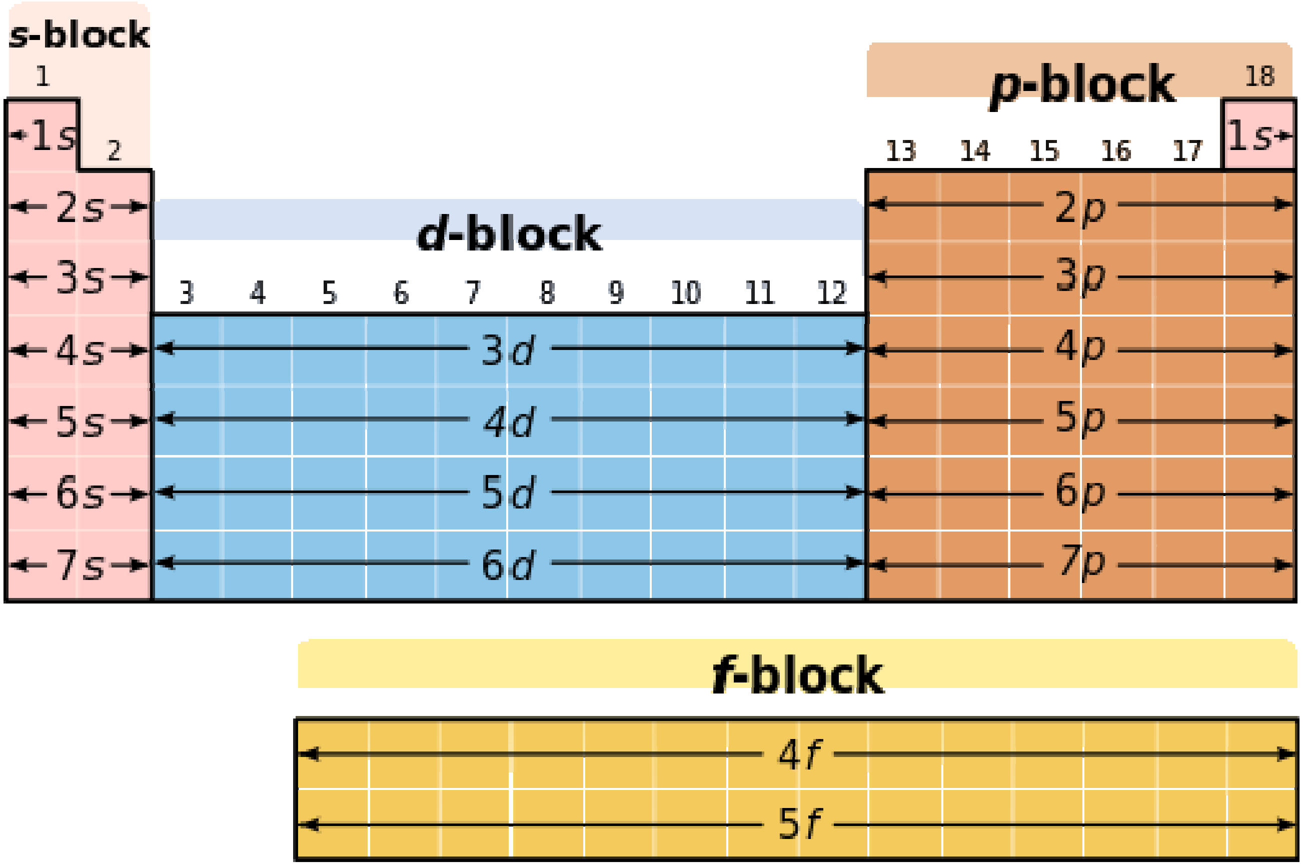
VI. Orbital diagrams of atoms in excited states

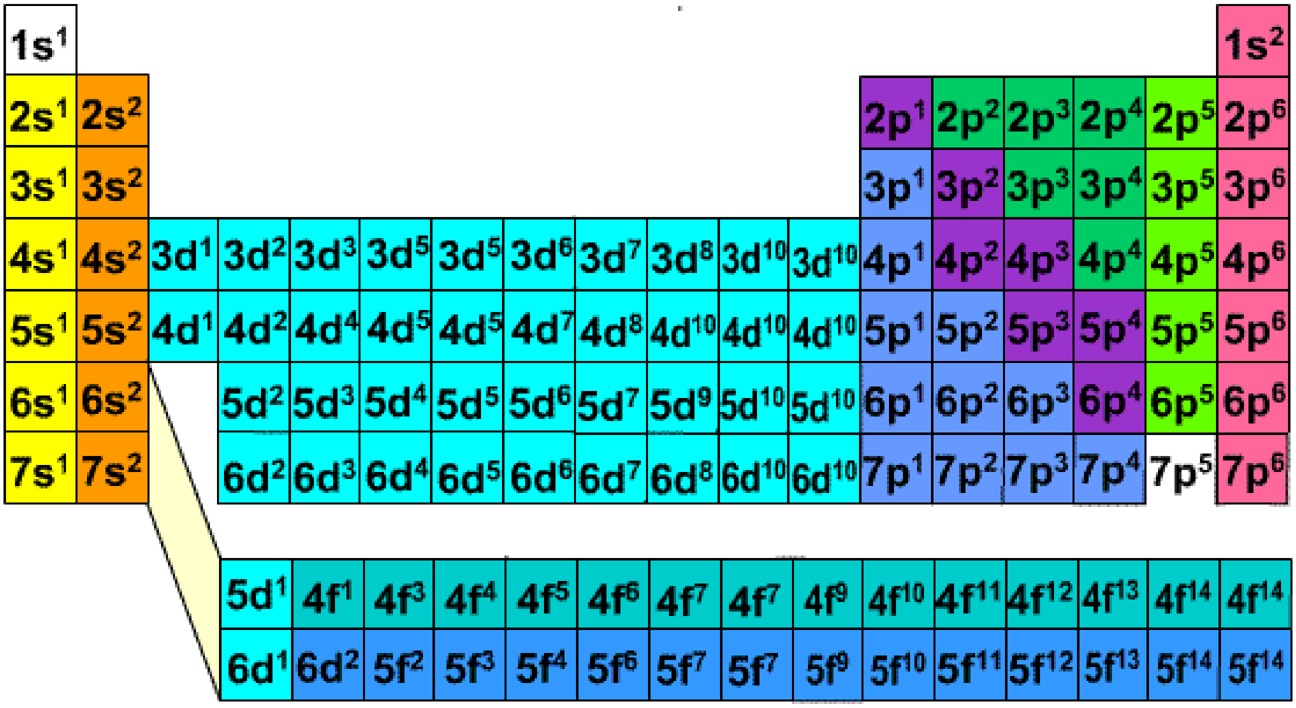
1. Sodium atom: _____ electrons

$1s^2 2s^2 2p^6 3s^1$

Ground state: 1s ☐ 2s ☐ 2p ☐☐☐ 3s ☐

One Possible Excited State: 1s ☐ 2s ☐ 2p ☐☐☐ 3s ☐ 3p ☐☐☐





$1s^1$																	$1s^2$
$2s^1$	$2s^2$											$2p^1$	$2p^2$	$2p^3$	$2p^4$	$2p^5$	$2p^6$
$3s^1$	$3s^2$											$3p^1$	$3p^2$	$3p^3$	$3p^4$	$3p^5$	$3p^6$
$4s^1$	$4s^2$	$3d^1$	$3d^2$	$3d^3$	$3d^5$	$3d^5$	$3d^6$	$3d^7$	$3d^8$	$3d^{10}$	$3d^{10}$	$4p^1$	$4p^2$	$4p^3$	$4p^4$	$4p^5$	$4p^6$
$5s^1$	$5s^2$	$4d^1$	$4d^2$	$4d^4$	$4d^5$	$4d^5$	$4d^7$	$4d^8$	$4d^{10}$	$4d^{10}$	$4d^{10}$	$5p^1$	$5p^2$	$5p^3$	$5p^4$	$5p^5$	$5p^6$
$6s^1$	$6s^2$		$5d^2$	$5d^3$	$5d^4$	$5d^5$	$5d^6$	$5d^7$	$5d^9$	$5d^{10}$	$5d^{10}$	$6p^1$	$6p^2$	$6p^3$	$6p^4$	$6p^5$	$6p^6$
$7s^1$	$7s^2$		$6d^2$	$6d^3$	$6d^4$	$6d^5$	$6d^6$	$6d^7$	$6d^8$	$6d^{10}$	$6d^{10}$	$7p^1$	$7p^2$	$7p^3$	$7p^4$	$7p^5$	$7p^6$

$5d^1$	$4f^1$	$4f^3$	$4f^4$	$4f^5$	$4f^6$	$4f^7$	$4f^7$	$4f^9$	$4f^{10}$	$4f^{11}$	$4f^{12}$	$4f^{13}$	$4f^{14}$	$4f^{14}$
$6d^1$	$6d^2$	$5f^2$	$5f^3$	$5f^4$	$5f^6$	$5f^7$	$5f^7$	$5f^9$	$5f^{10}$	$5f^{11}$	$5f^{12}$	$5f^{13}$	$5f^{14}$	$5f^{14}$

Name: _____________________ Per: ____

Note Questions: Modern Atomic Theory

Notes #8: Electrons in the Atom

1. Draw Rutherford's early model of the atom:
2. Rutherford's early model was wrong, and was replaced by the Bohr model. Explain what observations proved that the Rutherford model was incorrect:
3. Draw a diagram of the Bohr model:
4. What does "quantum" mean?
5. Explain why "quantum" energy levels are like a ladder:
6. What is an "excited state" for an electron?
7. What is a "ground state" for an electron?
8. What causes an electron to rise to an excited state?
9. What happens when an electron falls back to a ground state?
10. Why do different elements glow different colors when energized with heat or electricity?
11. Are orbitals and orbits the same thing?
12. What is an orbital?

Notes #9: Electron Orbitals

13. The principal ______________________ is related to the probable distance of the electron from the nucleus
14. The higher the energy level, the (closer/farther) an electron is from the nucleus (circle the correct answer)
15. Higher-energy orbitals can carry (more/fewer) electrons (circle the correct answer)
16. What formula is used to calculate the total number of electrons in an energy level?

Be sure to complete all questions on both sides of this paper!

17. SHOW YOUR WORK for how to Calculate the number of electrons in energy level 4 (your answer should be 32, NO WORK = NO CREDIT for the entire worksheet!!)

18. SHOW YOUR WORK for how to Calculate the number of electrons in energy level 5 (NO WORK = NO CREDIT for the entire worksheet!!)

19. S orbitals are found at energy levels ____ and above, are [insert shape(s) here] ____________________ and contain up to _____ electrons in total

20. P orbitals are found at energy levels ____ and above, are [insert shape(s) here] ____________________ and contain up to _____ electrons in total

21. D orbitals are found at energy levels ____ and above, are [insert shape(s) here] ____________________ and contain up to _____ electrons in total

22. F orbitals are found at energy levels ____ and above, are [insert shape(s) here] ____________________ and contain up to _____ electrons in total

Name: _________________________ Per: ____

Worksheet: Simple Electron Configurations

Write the full electron configurations of the following elements:
1) hydrogen: ____

2) helium: ____

3) aluminum: ____

4) fluorine: ____

5) potassium: ____

6) bromine: ____

7) vanadium (V, element #23): ____

8) bismuth (Bi, element #83): ____

The following electron configurations belong to which elements? (assume that they are neutral atoms):
9) $1s^2$ ____

10) $1s^22s^1$ ____

11) $1s^22s^22p^63s^23p^4$ ____

12) $1s^22s^22p^63s^23p^64s^1$ ______________________________________

13) $1s^22s^22p^63s^23p^64s^23d^7$ ___________________________________

14) $1s^22s^22p^63s^23p^64s^23d^{10}4p^5$ _____________________________

15) $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^8$ ___________________

16) $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}6p^67s^2$ ______________

Read these directions carefully and fully!
Determine if the following electron configurations are correct or not. (assume that they are neutral atoms). Write "correct" if it is right, and write the element name next to it to prove that you read these directions. If incorrect, circle the incorrect part and then write a corrected version in the corresponding blank. If the configuration is otherwise correct but does not match the element, then alter the configuration to produce a correct match for the element listed.

17) For Neon: $1s^22s^22p^2$ ____

18) For Silicon: $1s^22s^23s^23p^2$ ____

19) For Zirconium: $1s^22s^32p^63s^23p^64s^24d^{10}4p^65s^25d^2$ ____

20) For Magnesium: $1s^22s^22p^63s^2$ ____

21) For Boron: $1s^22s^3$ ____

Name: ___________________________ Per: ____

Worksheet: Electron Configurations (With Ions)

Write the unabbreviated electron configurations of the following elements:
1) sulfur: ____

2) strontium: ____

3) aluminum: ____

4) manganese: ____

5) calcium ion: ____

6) chloride ion: ____

7) argon: ____

Write the abbreviated electron configurations (i.e. noble gas configurations) for the following elements:
8) titanium: ____

9) fluorine: ____

10) astatine: ____

11) beryllium: ____

12) cesium: ____

13) iridium: ____

14) rubidium ion: ____

15) krypton: ____

16) The electron configuration of ions that are not transition metals is (different from/identical to) the nearest noble gas. (circle one)

The following electron configurations belong to which elements (assume that they are neutral atoms):
17) $1s^2 2s^2 2p^6 3s^1$ ____

18) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^6$ _________________________________

19) [Kr] $5s^2 4d^{10}$ ____

20) [Xe] $6s^2 4f^{14} 5d^{10} 6p^2$ ____________________________________

21) [Rn] $7s^2 5f^{14} 6d^4$ ____

22. List at least five ions that have the same electron configuration as argon:
a. ________ b. ________ c. ________ d. ________ e. _________

Determine if the following electron configurations are correct or not. (assume that they are neutral atoms). Write "correct" if it is right, and identify the element. If incorrect, circle the incorrect part that is wrong and write a corrected version in the blank.
23) [Rn] $7s^2 5f^9 6d^2$ ____

24) [Ar] $5s^2 4d^{10} 5p^5$ ____

25) [Xe] $6s^2 4f^{10}$ ____

26) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4d^{10} 4p^6 5s^1$ ___________________________________

27) $1s^2 2s^2 2p^6 3s^3$ ____

28) [Xe] ____

Name: ______________________ Per: ____

Worksheet: Orbital Diagrams

A. Draw an orbital diagram for each of the following elements.

	Element	Orbital Diagram	# of valence electrons	# of core electrons
1	He			
2	Al			
3	Mn^{3+}			
4	C			
5	Ar			
6	N^{3-}			
7	Na^{+}			
8	S^{2-}			
9	Ne			
10	Ca^{+2}			

B. Draw an orbital diagram for each of the following elements using short hand notation.

	Element	Orbital Diagram	# of valence electrons	# of core electrons
1	Mn			
2	I			
3	Ba			
4	Fe			
5	Po			

Name: _________________________ Per: ____

Electron Configuration and Orbital Diagram Review and Practice

1. Electron configurations review: Write how many electrons can fit into each energy sublevel:
   s: ___________ p: ___________ d: ___________ f: ___________
2. When an element loses electrons, does it become positive or negative? ______
3. When an element gains electrons, does it become positive or negative? ______
4. Write the full electron configurations for the following elements, which are neutral unless otherwise specified:
   a. Beryllium: ______________________________________
   b. Oxygen: ____
   c. Potassium ion: ___________________________________
   d. Bromide ion: _____________________________________
   e. Nitride ion: _____________________________________
   f. Neon: ____
5. Write the abbreviated electron configurations for the following elements, which are neutral unless otherwise specified:
   a. Lithium: _______________________________________
   b. Lead ion with a charge of +4: ____
   c. Sodium: ____
   d. Argon: ____
   e. Iodide ion: ______________________________________
   f. Barium ion: _____________________________________
   g. Copper ion with a charge of +1: ____
6. When subtracting electrons, do you remove them from the orbitals with the highest or lowest principal quantum numbers? _________
7. Draw the full orbital diagram for the following elements, which are neutral unless otherwise specified:

Helium	
Sulfur	
Neutral iron	
Lithium ion	
Fluoride ion	

Be sure to complete all questions on the back of this page as well!

8. Draw the abbreviated orbital diagram for the following elements, which are neutral unless otherwise specified:

Sodium	
Neon	
Calcium	
Bromide ion	
Iron with a charge of +3	

Topic List: Modern Atomic Theory (Unit 4)

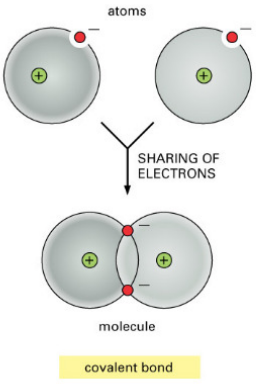
The following is a list of topics, concepts, and/or specific tasks that you are expected to have mastered. DO NOT RELY SOLELY ON THIS DOCUMENT!! It is NOT a study guide, nor is it in any way a substitute for the study guide that you should be making for yourself. There may be topics included on the exam that are not on this list. Use this list as a general starting place in your studying for the exam.

Things you should know or be able to do:

- Know how to calculate the number of electrons in an energy level
- Explain what a line spectrum emission is and why an atom produces one
- Explain how a line spectrum emission demonstrates the existence of quantized energy levels in an atom
- Explain why different atoms produce different line spectra
- Know the terms "excited state" and "ground state", what is needed for electrons to move from one to the other, and what is given off or absorbed when this happens
- For flame tests, know what causes the sample to give off light, in terms of where the light comes from and how the process occurs. You will probably need to study the textbook, notes and the flame test lab introduction, which is available on the class website.
- Describe Rutherford's early model of the atom
- Describe how we know that Rutherford's early model of the atom was wrong
- Describe the Bohr model of the atom and how we know that electrons have different energy levels
- What "quantum" means
- Explain what parts of the Bohr model tuned out to be correct, and which were wrong
- Describe the wave mechanical model of the atom
- Understand the terms "valence electron" and "core electron"
- Predict how many valence electrons an element or ions has based on its location in the periodic table and charge (if it's an ion)
- Count how many valence electrons an atom/ion has based on its electron configuration or diagram, and which kind of element is the only kind where neutral atoms can have a full energy level of electrons
- Know the difference between a sublevel and an orbital
- Know the following about sublevels
    - What they are
    - The different types
    - How many orbitals they can hold
    - How many electrons they can hold
    - What energy levels they can exist at (example: there is no "f" sublevel in energy levels 1, 2 or 3)
- Be able to explain what we know about the location and movement of electrons around an atom according to modern quantum theory
- Know how to use the periodic table to write/draw a full electron configuration or diagram for any atom or ion
- Know what electron "spin" is and how it is represented in an electron diagram
- Be able to write a full electron configuration OR a shortened one using noble gas notation for any atom or ion
- Be able to draw a full electron diagram that predicts one possible excited state for an electron
- Be able to predict which orbital is being filled at any point on the periodic table. Also be able to describe how electrons are being added to or taken away from it in order to turn any given atom into an ion.
- If given a list of sublevels (e.g. 1s, 4d, 6f, etc) be able to list them from lowest to highest energies

Unit 5: Chemical Bonding and Intermolecular Forces

I. Chemical Bonds



A. What is a chemical bond?

Answer: Chemical bonds are the ______________ holding two or more __________ together

1. ____________________ or ____________________ are formed when atoms bind together

B. Why do atoms bond together?

1. Answer: atoms __________ together because of electrostatic forces

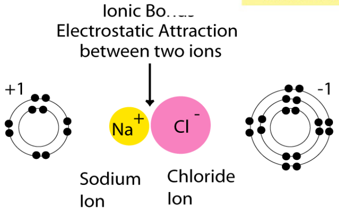
(_____________________ between ________________ and ________________ charges)

a. Examples:

i. Attraction between ______________________

ii. Attraction between ____________ of atom and ______________

of another _________



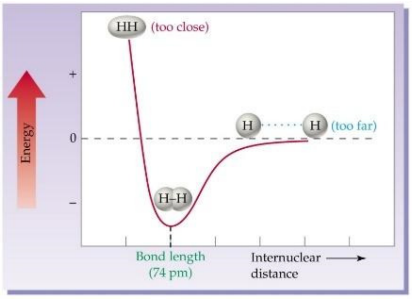
C. “Electrostatic forces” are also called ________________________________

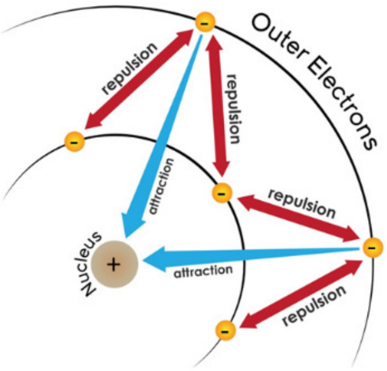
1. Opposite charges ______________ and like charges ____________

a. optimal bond length is the _____________ where attraction between ______________ and ________________ balances

the repulsion between ___________________

2. Bonds represent the state of lowest ________________________________ (maximum _____________________)





D. Bond energy

1. Energy required to break a bond

a. forming bonds _____________ energy

i. analogy: this is like allowing ___________ to snap together

b. breaking bonds _____________ energy

i. analogy: this is like the force required to pull ____________ apart

2. Higher bond energy = ______________________________

a. stronger ________________________ must be overcome in order to pull the _________ apart

i. analogy: strong bonds are like strong _______________: both require a lot of ___________ to pull them apart!

3. Lower bond energy = ______________________________

a. weaker ____________________________ make it easier to pull the _________ apart

i. analogy: weak bonds are like weak _____________: both require less _____________ to pull them apart

II. Types of Bonds

A. ___________ Bonding

1. One atom takes ____________________ from the other

2. ____________________ between oppositely ______________ ions

(________________________) holds them together

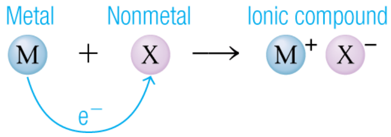
3. Results when ___________ react with ________________

4. Large difference in _______________________ Ex: __________

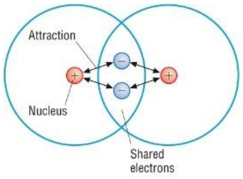
5. Electronegativity

a. Definition: tendency of an atom to attract _______________ toward itself

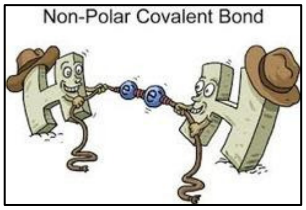
“JUST TAKE HIM! PLEASE!”
Na
Cl
“Don’t mind if I do.”

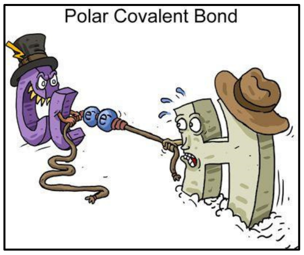


B. Covalent Bonding

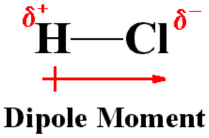


1. Involves sharing of ____________________ between _________________
2. Both atoms held together by ___________________ to the same ______________________
3. Two types: non-polar and polar
    a. Non-polar ________________________ bonding
        i. Equal Sharing of _____________________ between __________________
        ii. Bonds formed between ________________ or between identical _______
        iii. Both atoms attract the same ______________
        iv. ___________ or _______ difference in electronegativity
            Ex: ______ and _______

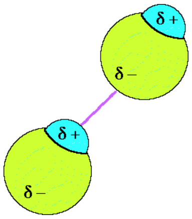




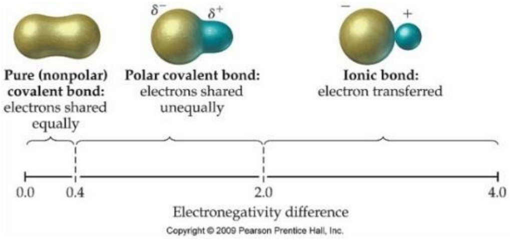
    C. ____________________________________ bonding
        i. ___________________ are shared _________________between atoms
        ii. One _______ attracts the shared _______________ much more
            _____________ than the other
            - Partial _____________________ on the atom with _______ electron density
            - Partial ____________________ on the atom with _______ electron density
        iii. Significant difference in ________________________, but not as much as
            in ______________________________
            Ex: _______ and ________



4. ___________________ moment
    a. When a __________________ has partial ___________ & __________ ends
    b. Results from polar covalent bonding
    c. Example: __________
        i. Water molecules _______________ oppositely-charged _________ of other ______________ molecules
        ii. Water molecules "_________" to each other strongly
        iii. Result: it takes a lot of ______________ to separate water molecules
            (high __________________, high surface _______________)



5. Use electronegativity differences to determine the type of ______________
    a. Difference of ______________ = non-polar covalent bond
    b. Difference of ______________ = polar covalent bond
    c. Difference of ______________ = ionic bond
    d. These are __________________________


Pure (nonpolar) covalent bond: electrons shared equally
δ− δ+
Polar covalent bond: electrons shared unequally
− +
Ionic bond: electron transferred
0.0
0.4
2.0
4.0
Electronegativity difference
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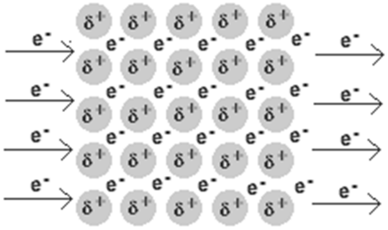
C. Metallic Bonding

1. Vacant ____ and _____ orbitals in outer energy levels ____________________ allowing outer (valence) electrons to move _________ throughout the metal 2. Metallic ______________ result from attraction between metal _____________ and the surrounding ____________ of ____________________

3. Metals are good __________________________________ because their
    _______________________ can move _____________
4. Metals are _______________________, malleable and ductile because the
    cations in the _________ of electrons are able to _______________ around
    and not locked into place.

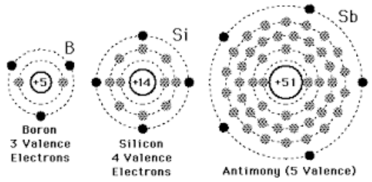


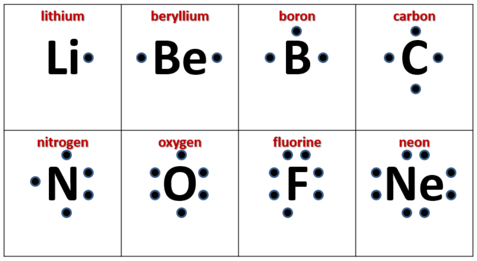
I. Valence electrons

A. The electrons in the ______________________ energy level

1. Available for ____

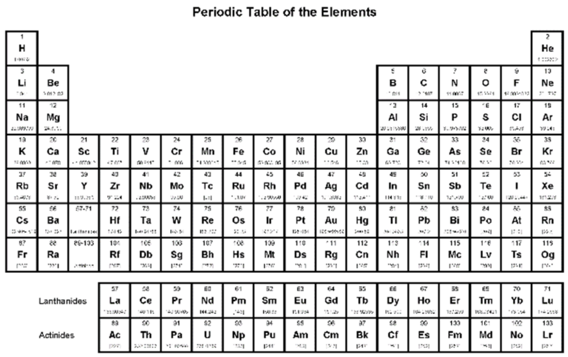
2. All other electrons are _________________________________





Label how many valence electrons in each group below:

Periodic Table of the Elements



H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	57-71	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	89-103	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og

Lanthanides	La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Actinides	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

II. Dot Notations

A. An atom's ____________ electrons can be represented

by Lewis dot notations

III. Showing bonds with Lewis dot structures

A. Lewis Dot Structures show how ____________________ are shared in a molecule's __________________ bonds

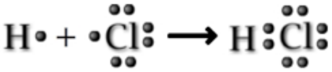
1. Lewis dot structures are NOT used for ______________ bonding because ________ don't ____________ electrons

B. All ________________ electrons are shown in Lewis dot structures, even the ones not _______________

as part of the ________________________

A covalent bond can be illustrated with

hydrogen chloride: (AKA hydrochloric acid)



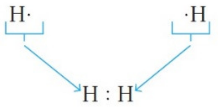
The hydrogen atom shares its valence electron with chlorine to form a covalent bond.

Drawing Lewis dot structures can help in understanding chemical bonding.

IV. Duet and octet rules

A. Duet rule

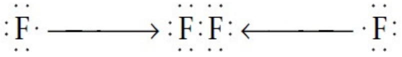
1. Elements #1-5 are most stable when they have _______ valence electrons



Helium already has 2 electrons, so it doesn't bond with anything

B. Octet rule

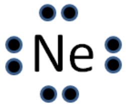
1. Many elements from #6 and up are most stable when they have ____ valence electrons



F atom with seven valence electrons

F_2 molecule

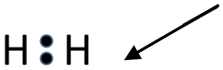
F atom with seven valence electrons



Neon already has 8 electrons, so it doesn't need to bond with anything

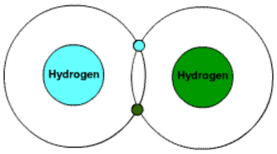
C. Example #1: Hydrogen

has only one electron, 1st energy level can hold 2

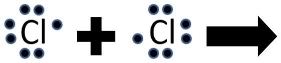


Two hydrogens share their electrons. Each "feels like" it has a full 1st energy level

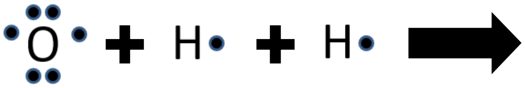
This is why hydrogen exists only as H_2 and never as H



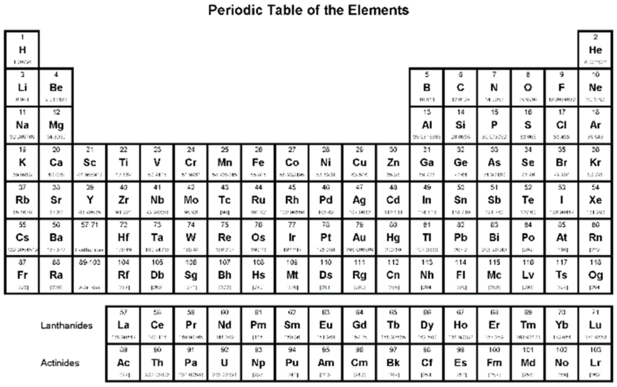
Example #2: Chlorine



Example #3: Water



Groups of the periodic table tend to form a certain number of chemical bonds (label how many for each group in the periodic table at right)


Periodic Table of the Elements
Lanthanides
Actinides

C. Two acceptable ways to draw Lewis structures:

1. With dots only (dots represent ___________________):

2. With lines (lines represent _________________ involved in ___________________)

V. Drawing Lewis Structures

A. The four steps of building a Lewis structure correctly that you need to memorize:

1. Count how many _______________________________
2. Build the ________________________
3. Add ____________________ until you have the correct total number
4. Check against the ___________ and ___________ rules

A. Example: Ammonia (NH_3)

1. Count the ________________________

a. N has _____, H has _____. ________________ = ______ total

2. Draw the molecule. Remember that each bond = _____ electrons.

3. Add ______________ until you have the correct total number (_____ in this case). That means add ____ more since you already have ______.

4. Check: does the molecule violate the octet or duet rules?

5. This molecule has the ______________ total number of electrons and the _________ and ___________ rules are satisfied. It is a valid structure.

NOTE: carbon will have _____ bonds and ______ lone pairs wherever possible

Practice examples can be done on a separate paper.

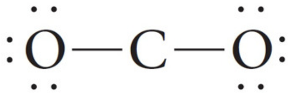
VI. Multiple Bonds

A. Consider the following problem

1. Carbon dioxide: ________________________ valence e^-

2. Is it a valid structure?

- there are 16 electrons just like there should be
- problem: not all atoms have a full octet.
- solution: __________________ bonds



Draw an arrow to indicate which atom in the structure above violates the octet rule

3. Make sure the structure is valid by checking the octet/duet rules and that it has 16 total electrons like it should (draw the correct structure below)

B. Step-by-step guide to drawing Lewis structures for multiple-bonded molecules:

1. Follow the four steps of building a Lewis structure first. DO NOT begin with _____________ or ____________ bonds!

2. Only consider multiple bonds if the correct number of valence electrons isn’t enough to satisfy the __________ and _________ rules for all the atoms

3. Move around ___________________________ pairs of _____________ to create multiple bonds until the _________ and _________ rules are satisfied

- DO NOT EVER add or subtract ___________ while attempting to do this

C. Definitions

1. Single bond: atoms share ____ pair of electrons

2. Double bond: atoms share _____ pairs of e^-

3. Triple bond: atoms share _____ pairs of e^-

a. Example of triple bond: N_2 (draw the process at right!)

VII. Polyatomic Ions

A. Consider the ammonium ion, NH_4^+

Step 1: count how many valence electrons:

a. N has 5, H has 1. 5+1+1+1+1 = 9 e- total

b. the total charge is +1. This means ___________________ one electron: ____________________________

Step 2: assemble the molecule:

Step 3: Check that you have the correct total number of ________________________

Step 4: Check against the ____________ and _______________ rules

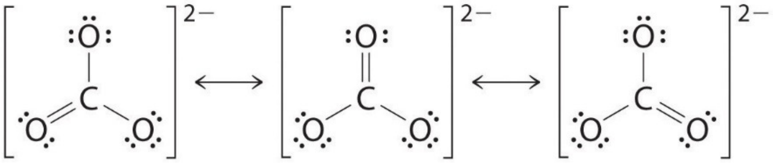
Step 5: Add ___________ and _______________:

VIII. Resonance structures

A. Sometimes more than ______ Lewis structure is valid for a molecule.

1. If there is more than _______ valid structure, the molecule will exist in multiple forms simultaneously

Example: carbonate ion ($CO_3^{2-)}$)



Now draw the ozone example below:

I. VSEPR Model

A. ____

1. The structure around a given atom is determined principally by minimizing _________________ between _________________

B. Lewis dot structures DO NOT accurately represent the __________ of real, three-dimensional molecules

C. Steps to determining molecular structure

1. Draw _______________________

2. Count number of __________________

3. Determine how to place ________________ as far ___________ as possible

4. ___________________________________

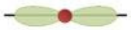
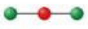
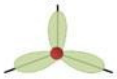
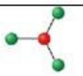
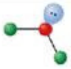
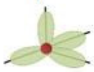
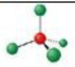
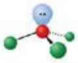
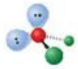
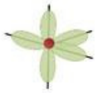
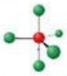
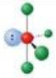
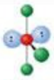
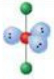
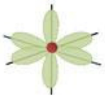
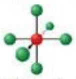
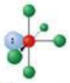
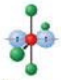
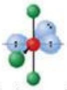
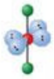
D. Molecular vs electron geometry

1. ____________________ geometry: include both ______________ AND __________________________ when determining shape

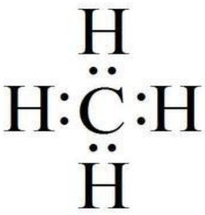
2. ____________________ geometry: look at the ______________ only when determining the _______________

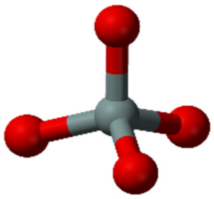
a. Note that while electrons do _____________ the ___________________ geometry, they

_______________ counted as a part of it.

Number of Electron Dense Areas	Electron-Pair Geometry	Molecular Geometry				
		No Lone Pairs	1 lone Pair	2 lone Pairs	3 lone Pairs	4 lone Pairs
2 	Linear	 Linear				
3 	Trigonal planar	 Trigonal planar	 Bent			
4 	Tetrahedral	 Tetrahedral	 Trigonal pyramidal	 Bent		
5 	Trigonal bipyramidal	 Trigonal bipyramidal	 Sawhorse	 T-shaped	 Linear	
6 	Octahedral	 Octahedral	 Square pyramidal	 Square planar	 T-shaped	 Linear

Example #1: Structure Determination of Methane by VSEPR

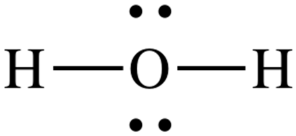




Methane = CH_4

- Build _________________________ first
- Count the domains (there are ______)
- Electron geometry: _________________________
- Molecular geometry: __________________________

Example #2: Structure Determination of Water by VSEPR

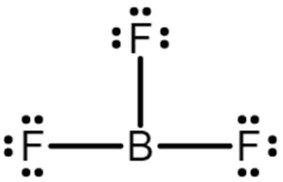


Water = H_2O

- Build _________________________ first
- Count the domains (there are ______)
- Electron geometry: _________________________ - Molecular geometry: __________________________

Example #3: Structure Determination of BF_3 by VSEPR

- Build __________________________ first
- Count the domains (there are _______ because boron does not follow the ______________ rule)
- Electron geometry: __________________________
- Molecular geometry: ___________________________

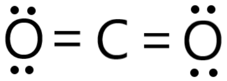


III. Molecules with ____________________ bonds

A. VSEPR Applies to __________________-bonded molecules as well

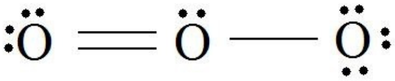
B. Examples with _______________ bonds

1. Ex. #1: Carbon Dioxide
   a. Central atom has ____ domains
   b. Electron geometry = ____________
   c. Molecular geometry = ____________



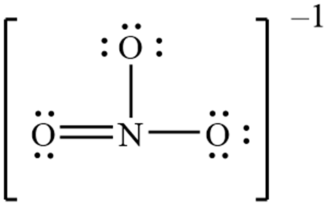
$O=C=O$

2. Ex. #2: Ozone
   a. Central atom has ____ domains
   b. Electron geometry = ____________
   c. Molecular geometry = ____________



$O=O-O$

3. Ex. #3: Nitrate Ion
   a. Central atom has ____ domains
   b. Electron geometry = ____________
   c. Molecular geometry = ____________



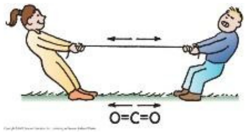
C. Triple-bonded molecule example: ethyne

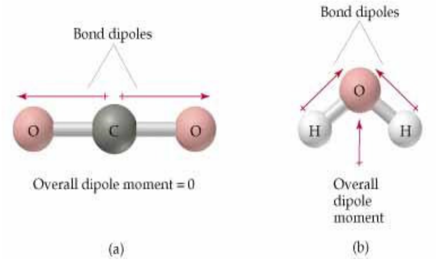
1. Central atoms both have ____ domains
2. Electron geometry = ____________________
3. Molecular geometry = ____________________

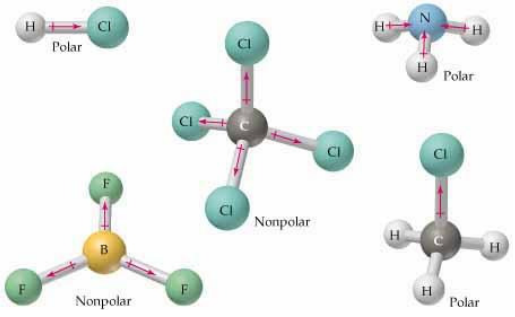
$H-C\equiv C-H$

IV. Using Shape to Predict Polarity of Molecules

1. Molecules with polar bonds can sometimes be __________________
2. If the shape is right, polar bonds can _____________ and _________________
   a. This happens in linear, trigonal planar, and tetrahedral shapes only, and ONLY if all bonds are ____________________







I. Many of the _________________ of a substance are determined by the balance of _____________ acting between its ___________

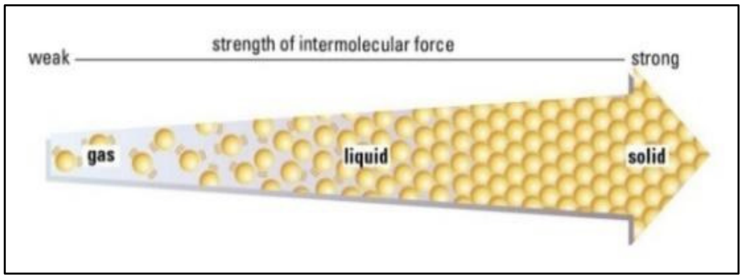
A. Forces pulling the particles apart

1. __________________ of the particles

a. Energy is _______________ at high temperatures

i. __________________ energy = more likely to be in a ________________

ii. (this is why things tend to _____________ at ___________ temperatures)



B. Forces pulling the particles together

1. Intermolecular forces

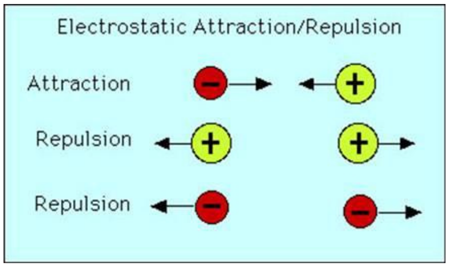
a. refers to electrostatic _________________ between particles

i. Stronger intermolecular forces = ___________ likely to be a __________ at a given temperature (particles ______________)

- higher _______________ and __________________ points

ii. Weaker intermolecular forces = ___________ likely to be _________ at a given temperature (attraction is too ___________ to ________ together)

- lower _______________ and __________________ points



II. Intermolecular Forces

A. Electrostatic ______________ between atoms or ___________________

1. Not nearly as strong as intramolecular bonds (AKA ______________ bonds)

2. Responsible for holding the particles ______________ in ________________ and _________________

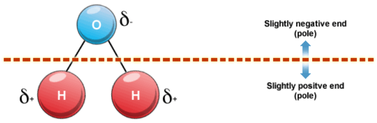
a. Also present in _________, but too _________ to hold the particles together

3. Strongly influence physical properties such as __________ points, ____________ points and _____________ pressures

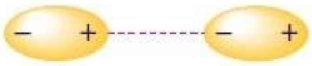
B. Types of Intermolecular Forces

1. ______________________ interactions

a. _______________ end of one polar molecule is attracted to the _______________ end of another



b. Example: ___________ molecules are _____________ and ______________ together strongly

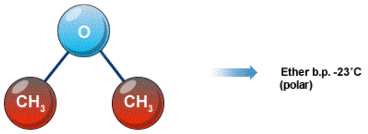


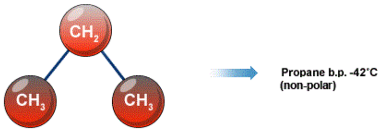
c. Dipole-dipole interactions become weaker in the gas phase due to large distances between individual molecules

d. Effect of dipole-dipole interactions:

higher _____________ and _____________points compared to similar _____________________ molecules

i. more ______________ needed to pull the _______________ apart means ______________ temperatures for melting and boiling





2. ____________________ Bonding

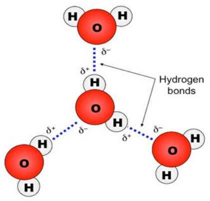
a. An especially __________________ form of ___________________ interaction

b. Occurs in molecules where __________________ is bound to an _______________ element

(e.g. _______________, nitrogen, fluorine)

c. _____________ attraction means that a lot of ______________ is required to pull the molecules away from each other

i. Result: _______________ melting and boiling points

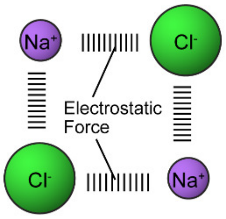


3. Ion-Ion Forces

a. Attraction between ____

b. The ____________________ type of intermolecular force

i. result: _______________________ melting and _________________ points



4. Ion-Dipole Forces

a. Attraction between ______________ and _____________________________

b. Relatively ______________ force of attraction

c. Make it possible for _____________ substances to ________________ in ______________ solvents.

(e.g. _______ dissolving in water)

5. London Dispersion Forces

a. ________________ force of __________________ between _____________________ molecules or individual atoms

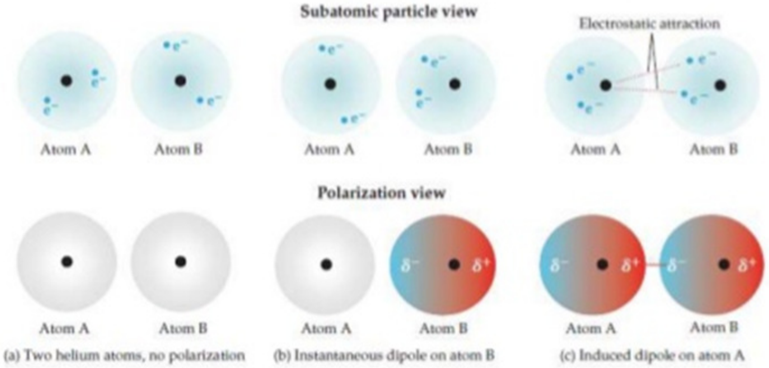
b. How London dispersion forces form:

i. ____________________ move _________________ around atom

ii. Sometimes they happen to ______________ up on one _________

iii. If the electrons group on one _______________, they form a temporary, very weak __________________

iv. This induces (______________) a temporary _______________ in other atoms/molecules and they are _______________ together



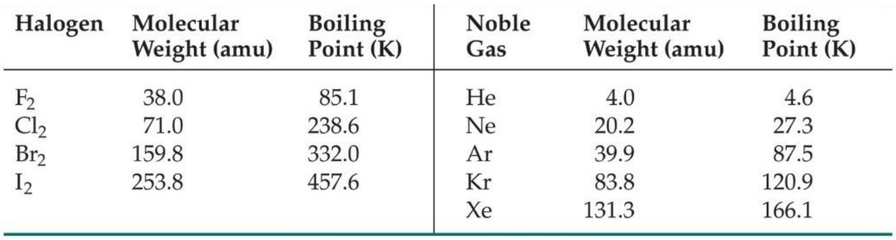
(a) Two helium atoms, no polarization
(b) Instantaneous dipole on atom B
(c) Induced dipole on atom A

Dispersion forces. “Snapshots” of the charge distribution for a pair of helium atoms at three instants.

c. London dispersion forces become stronger in _________________ atoms and molecules

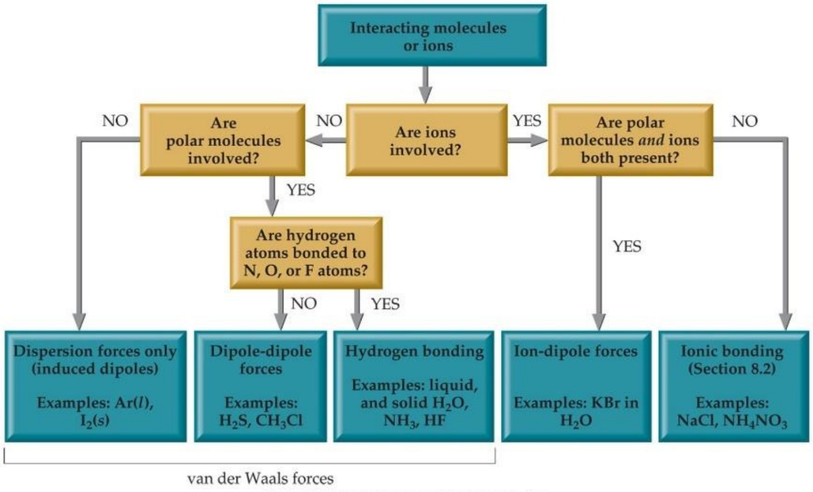
i. _______________ electron clouds are easier to ______________________

ii. This is why large ________________________ molecules can have very high melting and boiling points (e.g. ________, which is non-polar but has higher boiling and melting points than water due to the very _________________ size of its ________________________)



Halogen	Molecular Weight (amu)	Boiling Point (K)	Noble Gas	Molecular Weight (amu)	Boiling Point (K)
F_2	38.0	85.1	He	4.0	4.6
Cl_2	71.0	238.6	Ne	20.2	27.3
Br_2	159.8	332.0	Ar	39.9	87.5
I_2	253.8	457.6	Kr	83.8	120.9
			Xe	131.3	166.1

Summarizing Intermolecular Forces

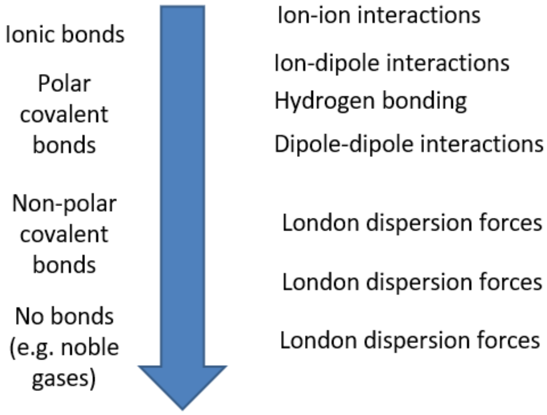

Interacting molecules or ions
Are ions involved?
NO
Are polar molecules involved?
NO
YES
Are hydrogen atoms bonded to N, O, or F atoms?
NO
YES
YES
Are polar molecules *and* ions both present?
YES
NO
Dispersion forces only (induced dipoles) Examples: Ar(l), $I_2(s)$
Dipole-dipole forces Examples: H_2S, CH_3Cl
Hydrogen bonding Examples: liquid, and solid H_2O, NH_3, HF
Ion-dipole forces Examples: KBr in H_2O
Ionic bonding (Section 8.2) Examples: NaCl, NH_4NO_3
van der Waals forces

Reminder (very important concept):

Strong intermolecular forces = ________ boiling and melting point for a substance

Weak intermolecular forces = ________ boiling and melting points for a substance

_______________ interactions



______________ interactions

IV. Effects of Intermolecular Forces on Solubility

A. “Solubility” = the ability of one substance to ______________ in another

B. “Like dissolves __________”

1. Polar compounds can dissolve _____________ or ______________

a. examples:

- salt (ionic) dissolves in ____________ (polar)
- ethanol (polar) dissolves in _____________ (polar)
- _________ (non-polar) does NOT dissolve in _____________ (polar)

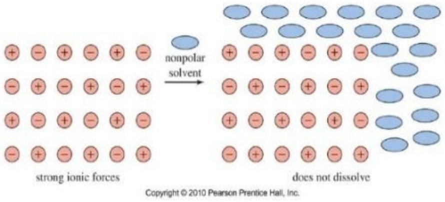
2. Non-polar substances can dissolve ONLY in _____________________

____

a. examples:

- ______ (non-polar) dissolves in ________ (Tetrachloroethylene, a non-polar substance)

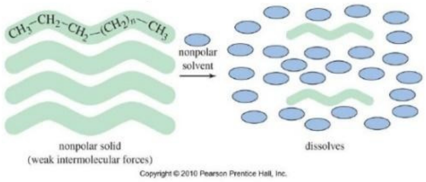
Polar Solute in Nonpolar Solvent



The solvent cannot break apart the intermolecular interaction of the solute, so the solid will not dissolve in the solvent.

Think of oil and water not mixing

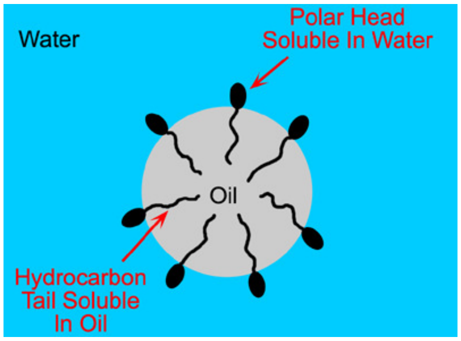
Nonpolar Solute with Nonpolar Solvent

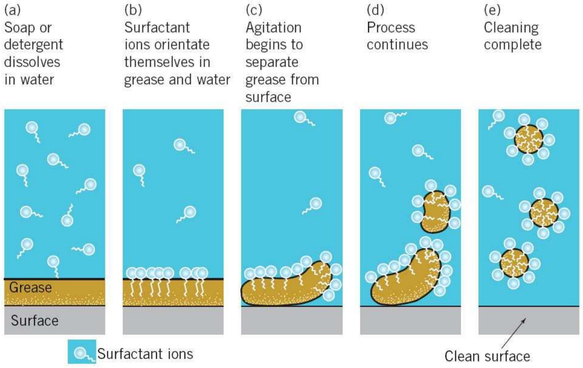

CH3–CH2–CH2–(CH2)n–CH3
nonpolar solvent
nonpolar solid (weak intermolecular forces)
dissolves
Copyright © 2010 Pearson Prentice Hall, Inc.

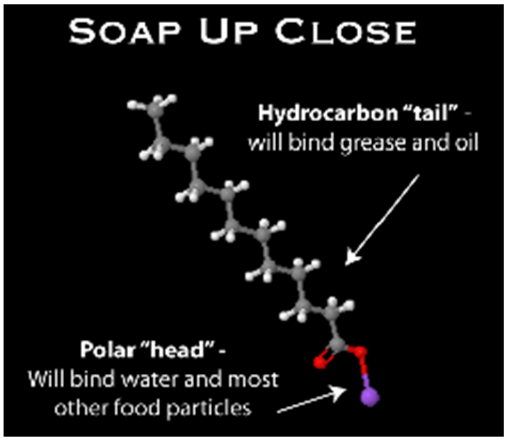
The weak intermolecular attractions of a nonpolar substance are overcome by the weak attractions for a nonpolar solvent. The nonpolar substance dissolves.

3. So how come you can use ________ and water (_______________) to wash away __________________ (non-polar)?

a. Answer: soap is an emulsifier, with a _____________ end that is attracted to the __________ molecules, and a ______________ end that is attracted to the ______




(a) Soap or detergent dissolves in water
(b) Surfactant ions orientate themselves in grease and water
(c) Agitation begins to separate grease from surface
(d) Process continues
(e) Cleaning complete
Grease
Surface
Surfactant ions
Clean surface

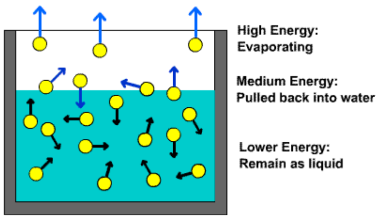

SOAP UP CLOSE
Hydrocarbon “tail” - will bind grease and oil
Polar “head” - Will bind water and most other food particles

I. Vaporization (AKA “Evaporation”)

A. Conversion of __________ into a _____

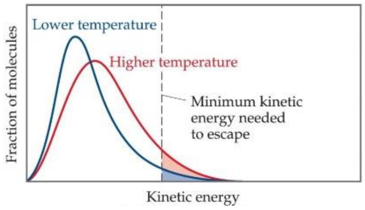
B. Only particles with enough kinetic _________________ to escape the _________________ forces of other particles can escape to the _______ phase

C. Particles with ____________ intermolecular forces evaporate more easily than particles with ____________ intermolecular forces


High Energy: Evaporating
Medium Energy: Pulled back into water
Lower Energy: Remain as liquid

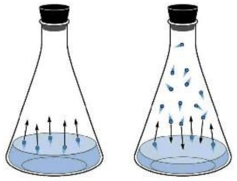
1. At any temperature, at least ________________ particles have enough ________________ to escape to the________ phase

2. At ________________ temperatures, a greater percentage of particles have enough ____________________ to escape



II. Vapor Pressure

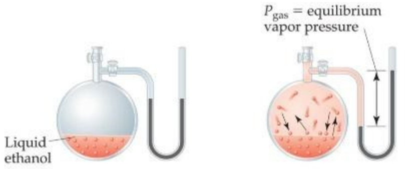
A. If a __________ is placed in a _______________ container, the particles will exert a _____________________ as they _________________________



B. As ________________ builds, it will reach a state of _____________ when it __________________ at the same rate that it ___________________

C. The _________________ exerted by the ____________ at equilibrium is _____________________________

D. When _________________ increases, _________________________ increases as more _____________ have enough _____________ to overcome attractive intermolecular forces and escape to the ____________ phase.



III. Boiling Point and Vapor Pressure

A. liquids ________ when vapor pressure = surrounding pressure

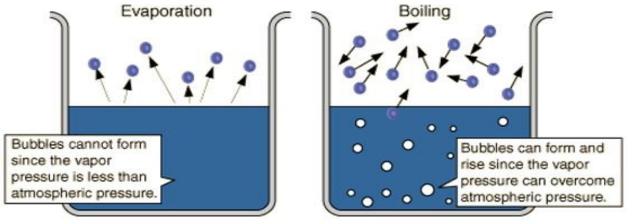
B. Vapor pressure and intermolecular forces

1. __________ intermolecular forces = ______ vapor pressure

a. _________ boiling point

2. ___________ intermolecular forces = ________ vapor pressure

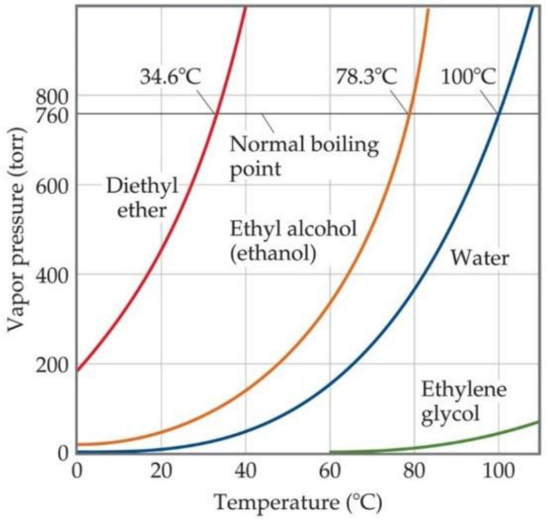
a. _________ boiling point


Evaporation
Boiling
Bubbles cannot form since the vapor pressure is less than atmospheric pressure.
Bubbles can form and rise since the vapor pressure can overcome atmospheric pressure.

C. High altitude areas have __________ air ____________, so water ___________ at lower _________________________

Note: you will be expected to…

- if given a certain pressure, be able to determine what the boiling temperature will be
- if given a certain boiling temperature, be able to determine what pressure will be needed to cause the substance to boil at that temperature



D. Volatility

1. Volatility is a measure of how rapidly a substance _________________

a. high __________________ = high _____________________

Substance evaporates _________________ due to ________________ intermolecular forces

b. low __________________ = low _______________________________

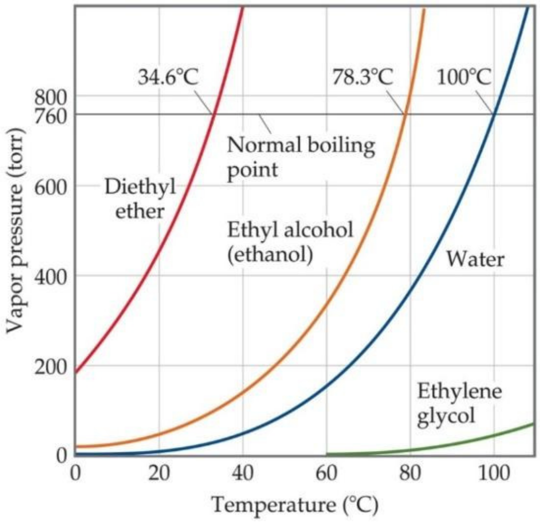
Substance evaporates _________________ due to ________________ intermolecular forces

E. Practice for how to use the graph at right

Use these videos to review (or to learn) how to complete the tasks below:

https://www.youtube.com/watch?v=VMOZh3_Lrdg

https://www.youtube.com/watch?v=5hUFw4su6fM



1. Determine the boiling point of diethyl ether at a pressure of 600 torr

2. Determine the boiling point of ethyl alcohol at a pressure of 400 torr

3. Determine the pressure needed to cause ethyl alcohol to boil at 50°C

4. Determine the pressure needed to cause water to boil at 80°C

5. Determine which substance has the lowest boiling point regardless of pressure

6. Determine which substance has the highest boiling point regardless of pressure

7. Determine which substance is the most volatile

8. Determine which substance is the least volatile

Name: ______________________ Per: ___

Comprehension Questions: Bonding and Intermolecular Forces

Notes #11 - Chemical Bonds

1. What is a chemical bond?
2. Why do atoms stick together and form chemical bonds?
3. What are electrostatic forces?
4. What are coulombic forces?
5. What is the relationship between bonds and potential energy?
6. Circle the correct answers: breaking bonds (absorbs/releases) energy, whereas forming chemical bonds (absorbs/releases) energy
7. Higher bond energy results in (weaker/stronger) bonds, and lower bond energy results in (weaker/stronger) bonds
8. Ionic bonding occurs between what two classes of elements?
9. Why do the atoms stick together in ionic bonding?
10. What is electronegativity?
11. Covalent bonds involve sharing what?
12. What is the difference between polar and non-polar covalent bonding? (*hint: your answer should involve electrons)
13. A molecule has a dipole moment when it has what?
14. List the approximate differences in electronegativity that produce the following bonds:
    a. Non-Polar Covalent: _________________
    b. Polar Covalent: _________________
    c. Ionic: _________________
15. Circle the correct answers: in metallic bonding, the nucleus (moves/stays put) and the electrons (move/stay put)

Notes #12 – Lewis Structures

16. What are valence electrons?
17. Valence electrons are available for what?
18. Are Lewis dot structures used for covalently-bonded compounds?
19. Are Lewis dot structures used for ionically-bonded compounds?
20. Do Lewis dot structures show electrons not part of a chemical bond?
21. Elements #1-5 are most stable with ______ electrons
22. Elements #6 and up are generally most stable with ______ electrons
23. What TWO things do you need to do to indicate electrical charge in a Lewis dot structure?
    a.
    b.

Notes #13: Shapes of Molecules

24. True or false: Lewis dot structures are accurate representations of the shape of a molecule

25. What do you look at when determining the electron geometry?

26. What do you look at when determining the molecular geometry?

27. Polar bonds can “balance” out to make non-polar molecules, but only if they have the right shape. List the three shapes below that make this possible:
    a.
    b.
    c.

Notes #14 - Intermolecular Forces

28. Many properties of a substance are determined by the forces acting on its particles.
    a. What forces act to pull the particles apart?

    b. What type of attraction acts to pull the particles together?

29. Does kinetic energy increase or decrease as temperature rises?

30. Are intermolecular forces stronger than or weaker than intramolecular forces (chemical bonds)?

31. Do strong intermolecular interactions increase or decrease boiling and melting points?

32. Is a material with high boiling point more likely or less likely to be a gas at room temperature?

33. Do weak intermolecular interactions increase or decrease boiling and melting points?

34. Is a material with low boiling point more likely or less likely to be a gas at room temperature?

35. List the three types of intermolecular forces that DO NOT involve ions, from strongest to weakest:

36. ________________ bonding is an especially strong form of dipole-dipole interaction

37. Are London dispersion forces most prominent in polar or non-polar atoms and molecules?

38. Are London dispersion forces strongest in larger or smaller atoms and molecules?

39. Polar substances can dissolve ___________ or ___________, whereas non polar substances can dissolve ______________

Notes #15: Vapor Pressure and Boiling Point

40. Particles with ______________ intermolecular forces evaporate more easily than particles with ___________ intermolecular forces

41. Liquids can evaporate at any temperature because at least some of the particles will have enough _________________ to escape into the vapor phase

42. What does it mean when vapor and liquid in a closed container reaches equilibrium?

43. When temperature _____________________, vapor pressure increases as more particles have enough energy to overcome ___________________ intermolecular forces and escape to the gas phase

44. What is the relationship between boiling point and vapor pressure?

45. Do **strong** intermolecular forces result in higher or lower vapor pressure?

46. Do **weak** intermolecular forces result in higher or lower vapor pressure?

47. What does it mean when a liquid is said to be *highly volatile*?

48. Circle the correct answers: highly volatile substances have (strong/weak) intermolecular forces, whereas low-volatility substances have (strong/weak) intermolecular forces

Name: _______________________ Per: _____

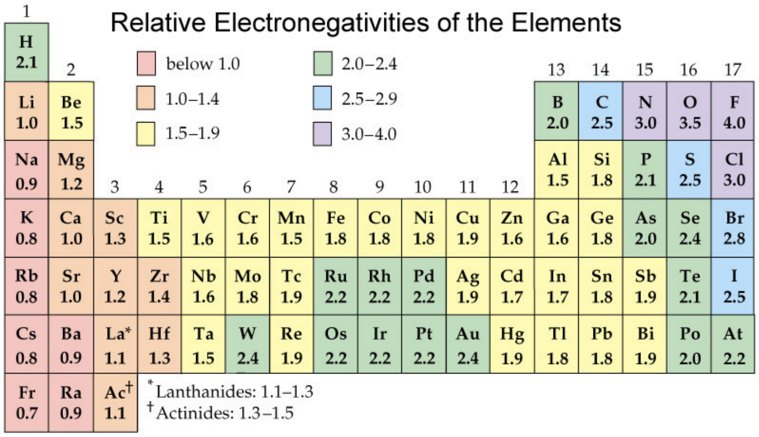
Worksheet: Predicting Chemical Bond Types

Review and Practice:

1. In ionic bonding, describe what happens to the electrons: ____
2. In non-polar covalent bonding, describe what happens to the electrons: ___________________________________
3. In polar covalent bonding, describe what happens to the electrons: _______________________________________

Define the following:

4. Chemical bond:
5. Electronegativity:



Relative Electronegativities of the Elements

🟥 below 1.0
🟧 1.0–1.4
🟨 1.5–1.9
🟩 2.0–2.4
🟦 2.5–2.9
🟪 3.0–4.0

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17
H 2.1																
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.8	Ba 0.9	La* 1.1	Hf 1.3	Ta 1.5	W 2.4	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	Ac† 1.1														

*Lanthanides: 1.1–1.3
†Actinides: 1.3–1.5

Explain, in your own words, how to use the periodic table above to predict bond type:

Practice predicting bond type. Important: show your work (see example) because YOU WILL GET ZERO CREDIT FOR THIS ENTIRE WORKSHEET IF YOU DON'T SHOW YOUR MATHEMATICAL WORK!!

You may use the following abbreviations: P. = polar, N.P. = non-polar, C = covalent. Do not abbreviate anything else.

C-H bond type: **P.C.**_ Your calculations here: **2.5 – 2.1 = 0.4**	CO_2 bond type: _________ Your calculations here:	Na-N bond type: ________ Your calculations here:	NO_2 bond type: _________ Your calculations here:
B-P bond type: _________ Your calculations here:	BH_3 bond type: _________ Your calculations here:	Zn-S bond type: _________ Your calculations here:	$FeCl_2$ bond type: ________ Your calculations here:
Fe-O bond type: ________ Your calculations here:	I_2 bond type: ___________ Your calculations here:		

Name: ______________________ Per____

Worksheet: Drawing Lewis Structures

Draw Lewis Structures for each of the following molecules. Don’t forget to include brackets charges for ions.
NOTE: Carbon is most stable when it has 4 bonds and no lone pairs of electrons!

Answer this before starting: carbon is most stable when it has ____ bonds and ____ lone pairs

1. CH_4	2. NH_3	3. H_2O
4. C_2H_2	5. $C_2O_4^{2-}$	6. HCN
7. NO_3^-	8. OH^-	9. NO^+
10. H_3O^+	11. NH_2^-	12. NH_4^+

Don’t forget to complete the problems on the other side of this page too!

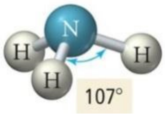
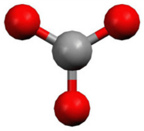
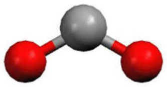
13. C_2H_4	14. H_3COH	15. H_2CO
16. PCl_3	17. CHCH	18. $SiCl_2Br_2$

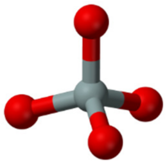
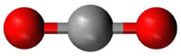
Ensure that all carbons in your Lewis structures have 4 bonds and no lone pairs!

Name: ______________________________ Per: ____

Molecular Geometry Practice Sheet

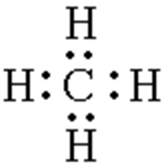
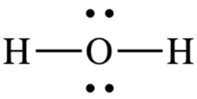
Part 1: Identifying Shapes

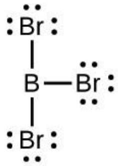
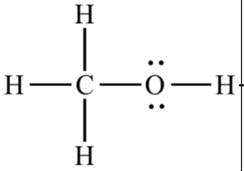
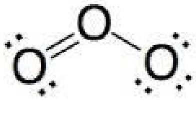
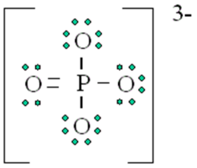
Picture	Name of the Molecular Shape
 (three atoms, all bent downward)	
 (three atoms, all flat)	
	

Picture	Name of the Molecular Shape
	
	
What type of geometry shows only the atoms and NOT the non-bonding electrons? ____________________	What type of geometry shows all bonding/non-bonding electrons? ____________________

Part 2: Practicing with Structures

Directions: Read these directions. If provided, use the modeling kit provided to build each molecule. Be sure to fill out all the blanks, and draw a smiley face next to your name for five points, and follow the style of the examples below. Lewis structure need to be neat and readable, or no credit. Use this worksheet as practice to help you to draw Lewis structures, as well as being able to use them to determine the shape of a molecule more easily.

Formula	Name of Molecule	Total Number of Valence Electrons	Lewis Structure	Electron Geometry	Molecular Geometry	Resonance or No Resonance ?	Polar or Non-Polar Bonds? (Show Work!)	Polar Molecule or Non-Polar Molecule?
CH_4	Methane	8		Tetrahedral	Tetrahedral	No Resonance	2.5 – 2.1 = 0.4 Polar bonds	Non-polar molecule
H_2O	Water	8		Tetrahedral	Bent	No Resonance	3.5 – 2.1 = 1.4 Polar bonds	Polar molecule
CCl_4	Carbon Tetrachloride	32					3.0 – 2.5 = 0.5 Polar bonds	
NO_3^-	Nitrate					Resonance	3.5 – 3.0 = 0.5 Polar bonds	

Formula	Name of Molecule	Number of Valence Electrons	Lewis Structure	Electron Geometry	Molecular Geometry	Resonance or No Resonance ?	Polar or Non-Polar Bonds? (Show Work!)	Polar Molecule or Non-Polar Molecule?
NH_4^+	Ammonium							
BBr_3	Boron Tribromide						2.1 – 2.0 = 0.1 Non-Polar bonds	Non-polar molecule
NF_3	Nitrogen Trifluoride							
H_2S	Hydrogen Sulfide							
C_2H_4	Ethene							
CH_3OH	Methanol			Oxygen Carbon	Oxygen Carbon		C-H C-O O-H	
CO_2	Carbon Dioxide							
O_3	Ozone							
PO_4^{3-}	Phosphate							

Name: ________________________ Per: ____

Worksheet: Vapor Pressure and Boiling Point

Part 1: Review and practice: circle the correct answers where appropriate.

1. STRONG intermolecular forces tend to produce (high/low) vapor pressure, (high/low) volatility, and (high/low) boiling points

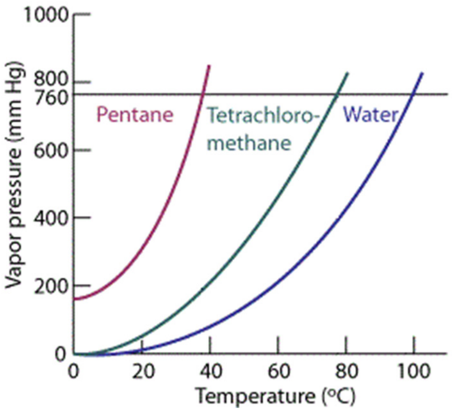
2. WEAK intermolecular forces tend to produce (high/low) vapor pressure, (high/low) volatility, and (high/low) boiling points

3. A substance with high vapor pressure tends to evaporate (quickly/slowly), whereas a substance with low vapor pressure tends to evaporate (quickly/slowly).

4. Define volatility:

5. Vapor pressure (increases/decreases) when temperature increases

6. Liquids boil when vapor pressure is equal to ____________________________



Part 2: Basic graph interpretation

1. What is the y-axis? ____

2. What is the x-axis? ____

3. As temperature rises, vapor pressure (rises/falls)

4. At any given temperature, which of the three substances has the highest vapor pressure? _______________ Which one has the lowest vapor pressure? _________________

5. Which of the three substances is the most volatile? ___________ Which one is the least volatile? _________________

6. Which substance has the strongest intermolecular forces? _____________ Which one has the weakest? _____________

Part 3: Practice using the graph

1. Determine the boiling point of pentane at a pressure of 760 mmHg _____

2. Determine the boiling point of tetrachloromethane at a pressure of 400 mmHg __________

3. Determine the boiling point of water at a pressure of 300 mmHg _______

4. At any given pressure, which substance has the highest boiling point?________ Does this substance has the strongest or weakest intermolecular forces? _______________

5. Determine the pressure needed to produce a boiling point of 40°C for tetrachloromethane: _______

6. Determine the pressure needed to produce a boiling point of 50°C for water: _______

7. Determine the pressure needed to produce a boiling point of 25°C for pentane: _______

8. Water boils at about 85°C in Cuzco, Peru, a city at about 4,000 meters (that's 13,000 feet to you Americans) of altitude. Determine the pressure up there: _______

Topic List: Exam Unit 5 - Bonding and Intermolecular Forces

The following is a list of topics, concepts, and/or specific tasks that you are expected to have mastered. DO NOT RELY SOLELY ON THIS DOCUMENT!! This is NOT a study guide, nor is it in any way a substitute for the study guide that you should be making for yourself. There may be topics included on the exam that are not on this list. Use this list as a general starting place in your studying for the exam.

Notes #5: Formulas of Ionic Compounds

- There will be a section on the exam where you will predict the formulas of ionic compounds based on their charges

Notes #11: Chemical Bonding

- Know what polar covalent, non-polar covalent, ionic and metallic bonding are. Know how electrons are shared in each (if, that is, they are even shared at all)
- Be able to define the term "dipole" and be able to explain why a polar molecule has a partial positive charge on one end and a partial negative charge on the other
- Know the symbols for partial positive charge and partial negative charge
- Based on electronegativity values, be able to calculate whether a bond is ionic, polar covalent, or non-polar covalent. You do not need to memorize specific numbers, they will be given to you on the exam
- Know what electronegativity is and how to use it to predict which atom in a polar molecule will have a partial positive charge and which will have a partial negative charge
- Be able to describe similarities and differences in the physical properties (hardness, melting point, etc) of ionic vs molecular compounds
- Describe what a chemical bond is (in terms of electrons) and how it causes atoms to stick together
- Be able to describe the relationship between the "sea of electrons" model and key properties of metals
- Be able to explain the relationship between potential energy and stability.

Notes #12: Lewis Structures

- Be able to draw Lewis structures for any compound, charged or neutral
- Be able to draw Lewis structures with resonance and to recognize situations where a resonance structure is needed
- Know what a valence electron is, how to determine the number of valence electrons in an element, and how to represent them in a Lewis dot structure
- Know what the duet and octet rules are. Know how they apply to Lewis structures and how they predict the number of bonds an element can form
- Be able to show (and explain) the difference between bonding electrons and non-bonding electrons (AKA lone pairs) in a Lewis structure
- Be able to show HOW you know the correct number of valence electrons to show in a Lewis Structure

Notes #13 Shapes of Molecules

- Know the various possible shapes predicted by VSEPR (meaning, of course, that you need to know what VSEPR is)
- Given a molecule with polar covalent bonds, be able to determine based on its shape whether or not the overall molecule is polar
- Using a Lewis structure, be able to predict a molecule's electron and molecular geometries
- Be able to predict whether a structure containing polar covalent bonds would be polar or non-polar overall based on its geometry (shape) and composition (whether the bonds are all the same or not)

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Notes #14: Intermolecular Forces

- Be able to describe the relationship between the strength of intermolecular attractive forces and melting/boiling point
- Know the types of intermolecular forces and how they work.
- Be able to organize the types of intermolecular forces from strongest and weakest.
- Be able to explain what a dipole is and why they tend to stick together.
- Explain why compounds with London dispersion can have high boiling points
- Understand why some things will dissolve in a given solvent (e.g. water), whereas others will not.
- Describe what is meant by the expression "like dissolves like"
- Be able to explain why soap can dissolve both oil (non-polar) and water (polar)

Notes #15: Vapor Pressure and Boiling Point

- Be able to read and understand a vapor pressure vs temperature diagram.
- Know how to determine boiling point at a given pressure using a diagram, or how to determine the pressure which would be needed to boil at a certain temperature. Also know how to use the diagram to determine how volatile something is.
- Know what "volatile" or "not volatile" means.
- Know the effect of intermolecular forces on boiling point, freezing point and volatility
- Know the meaning of "equilibrium"

Note: Your free-response section will consist of drawing Lewis structures and determining things such as their electron and molecular geometries, number of valence electrons, resonance, etc. See your molecular geometry practice sheet for an idea of what you will be asked to do. Also, you will be asked to draw a diagram explaining one (or all) of the concepts covered in the notes about intermolecular forces. Also, you will predict formulas of ionic compounds.